GENERAL
SCIENCE

BY

S. WHITEHOUSE
Cowley School, St. Helens

AND

A. W. PRITCHARD
Selhurst Grammar School for Boys,
Croydon

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PREFACE

MORE than a dozen years' experience in teaching General Science has convinced us that any course for the study of this subject must have two main objectives. It must provide the opportunity for the pupils to acquire a body of scientific knowledge so that they can study and analyse their environment, and it must, by continued example, develop in them a thorough appreciation of the importance and nature of the Scientific Method so that they are capable of applying it to new investigations in later life. The cause and prevention of disease, the study of nutrition, the significance of soil studies, the study of inheritance and evolution all demonstrate the need for treating General Science as a whole subject if our pupils are to appreciate to the full the interrelations between the parts of their environment and themselves. We believe that this side of science has particular value in helping to build a better society.

The criticism is often levelled at General Science that it is superficial and does not make scientists. It is this criticism which we have attempted to meet by providing a course which gives full accounts of investigations and discusses the reasons for carrying them out and the nature and limitations of the inferences that can be drawn from them.

The book is designed to cover a four or five years' course up to the standard of School Certificate. We have included many experiments, and have introduced them, not as arid academic examples of particular phenomena, but as essential steps in the development of our investigations so that, as far as possible, the pupils are guided in finding out for themselves. Instructions for the proper performance of every experiment are given. We believe that by doing so we serve two purposes: we provide laboratory directions for the pupils and also assist those teachers who, because of their lack of training in one branch of science, are not familiar with the conduct of experiments

in that branch. We make no apology for having included more experiments than can usually be carried out in the time allotted to science in the school curriculum, since by doing so we give an opportunity for variation in the course. Although the book meets the demands of a School Certificate course we have not hesitated to go outside the syllabus when an interesting link could be made.

Our main aim has been to write a book which treats General Science as a whole subject and, at the same time, so to stress the importance of accurate observation and logical discussion, that if the pupils specialize later in one particular branch they will be able, from their training in General Science, to approach new problems in a proper scientific manner. For those who do not become scientists the course provides a survey of their environment which will open their eyes to the relationships which exist between many phenomena in the world.

July 1947   S. W.
                                                                              A. W. P.

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PART II

X

CARBON AND ITS COMPOUNDS

CARBON

WE have already studied the burning of carbon in air and in oxygen. The element carbon exists in various more or less pure forms.

1. Wood charcoal

Expt. 136. Put a piece of wood at the bottom of a crucible; fill the crucible with sand, and heat, gently at first and then strongly. When the crucible is cool examine the contents. Compare the piece of charcoal obtained with a piece of wood similar to the one placed in the crucible. It is smaller in volume and lighter in weight.

It consists largely of the element carbon with small amounts of other substances. Charcoal is made on a large scale by burning wood in a limited supply of air. Logs of wood are stacked vertically round a little brush-wood and the whole covered with earth to exclude the air (Fig. 95). Small holes are left at the bottom and top to create a very slight current of air, otherwise the wood would not burn at all. The brush-wood is lighted, and the whole left smouldering for days. When the wood has stopped burning the earth is removed. The charcoal retains more or less the shape of the original wood, but is shrunken.

Charcoal which has been freshly heated has the power of absorbing gases. It is present in charcoal biscuits, used for removing excess of gases from the food canal. It is also packed round the tops of sewer ventilators to absorb offensive gases. These gases are oxidized by the oxygen of the air which is also absorbed. Gas masks are often packed with specially prepared and activated wood charcoal to absorb the poison gases.

Expt. 137. Fill a test-tube with ammonia gas (p. 253) and invert over mercury in a dish. Introduce a small piece of freshly heated

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charcoal under the mouth of the tube. It will quickly float to the top. The mercury will begin to rise in the tube showing that the gas is being absorbed.

Wood charcoal is also used as a fuel, and as a material for filtering.
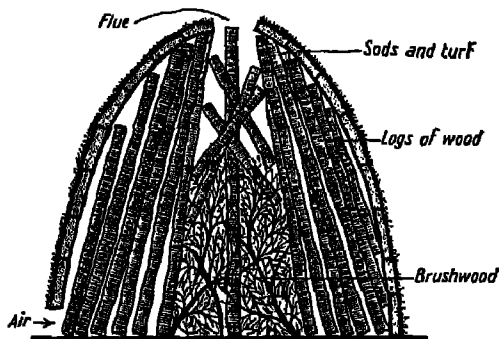
FIG. 95. An old method of charcoal burning still used in some districts where there is much wood.

2. Animal charcoal

This is made by heating bones in closed vessels. The resulting solid, when the liquid and gaseous products have been removed, is a porous mass of very finely divided charcoal. Only about 10 per cent. of this is pure carbon.

Animal charcoal has the peculiar property of removing the colour from certain solutions, e.g. litmus, and is used in industry for decolourizing brown sugar obtained from the sugar-cane and the sugar-beet. It is also used to filter well-water used for drinking purposes, as it filters off harmful material.

Expt. 138. Warm some diluted ink or litmus with freshly heated animal charcoal, and filter. The filtrate will be almost colourless.

3. Lamp-black

Expt. 139. Hold a piece of white porcelain (an evaporating dish) in the flame of a candle. Note the black deposit.

This variety of carbon is called lamp-black. It is made on a large scale by burning such substances as turpentine, petrol,

tar, fuel oils, in a limited supply of air. These substances are rich in carbon, and when burned under such conditions, the maximum of smoke is produced. This smoke, which consists of particles of carbon, is brought into contact with coarse cloth suspended in the burning chamber. The lamp-black collects on the cloth.

This variety of carbon is one of the purest, over 98 per cent. being pure carbon. Lamp-black has a wide use as a pigment in paint, shoe 'blacking', and printer's ink.

4. Coal and coke

These will be studied in detail later (p. 189).

5. Graphite

This is a crystalline form of carbon. All the forms mentioned above are amorphous (i.e. non-crystalline). Graphite is mined in many parts of the world. In Cumberland the mines, which were never large, are now almost exhausted. It is very soft, so soft, indeed, that it marks paper. It is used, mixed with finely-divided clay, in 'lead' pencils. At one time it was thought that there was lead in its composition and it was called 'black lead'. It is also used as a lubricant for machinery. Graphite is a good conductor of electricity and resists heat very well. For these reasons it is used in the manufacture of electrical apparatus. Mixed with water to form a paste, it is used as stove polish to prevent the iron from rusting.

6. Diamond

It is strange that this beautiful gem is made of the same element as graphite. It is the hardest substance known, yet graphite is so soft that it will mark paper. Graphite is black, yet the pure diamond is colourless. They are both crystalline forms. When mined, diamonds look like pieces of dirty, yellow glass. They are cut so as to give the greatest amount of 'sparkle' (Ch. XXXIII). On account of its hardness the diamond is used for cutting glass.

MOISSAN, a French chemist, succeeded in making small artificial diamonds. Pure charcoal, dissolved in molten iron,

was heated in an electric furnace at a temperature of 3,000° C. It was then suddenly cooled by plunging into molten lead. This subjected the carbon to an enormous pressure while cooling. When the mass of iron was dissolved away, small diamonds were found. They were of no commercial value however.

How do we know that all these different forms of carbon do consist of carbon? No one looking at a diamond would suspect that it was composed of the same element as lamp-black. If equal weights of these different forms of the same element are burned in oxygen, equal weights of carbon dioxide are produced. Not all the forms of carbon will burn easily in oxygen, but under certain conditions they can all be made to do so. The diamond can, under certain conditions, be changed to graphite. Graphite has been converted to very small diamonds as related above. (See p. 211 for allotropes of carbon.)

COMPOUNDS OF CARBON

There are two important oxides of carbon, one of which, carbon dioxide, CO_2, we have studied in connexion with the composition of the air, and the relations of animals and plants to the air. The other oxide is carbon monoxide, CO. This is a colourless, odourless, and poisonous gas which is important in modern life. We shall have to study this in some detail.

Carbon also combines with hydrogen to form *hydrocarbons*, some of which are very important, e.g. methane, or marsh gas, CH_4; acetylene, C_2H_2; benzene, C_6H_6; paraffin, turpentine, and petrol.

The compounds of carbon with hydrogen and oxygen, in which the last two elements are present in the proportion in which they occur in water, are called *carbohydrates*. We have already discussed two of them, sugar and starch.

In addition, carbon forms a large number of other compounds. All the living parts of plants and animals consist of carbon compounds, and all food taken by animals contains carbon compounds. The science of Organic Chemistry is devoted to the study of the compounds of carbon. It was once thought that the compounds of carbon occurring in living.

things could not be produced without the agency of life. Some
have now been prepared in the laboratory, but it is still true
that many of them have not yet been produced except by the
living plant-or animal. The name organic chemistry was given
at the time when it was unknown that some of these compounds
could be produced without living agents. We shall study some
of these organic compounds later (p. 286).

CARBON MONOXIDE
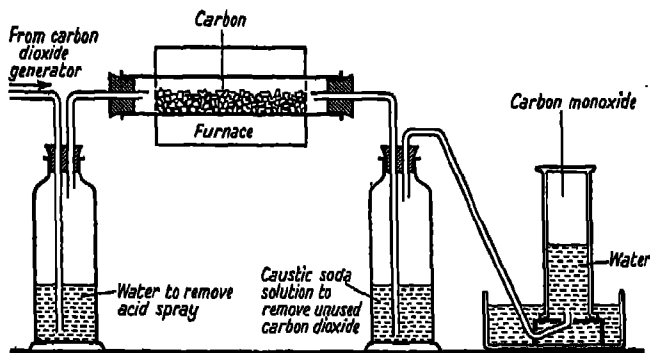
Fig. 96. Preparation of carbon monoxide.

When carbon dioxide is passed over strongly heated carbon,
carbon monoxide is produced.

$$CO_2 + C = 2CO.$$
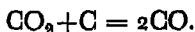
This is a reducing action, the carbon having taken oxygen
from the carbon dioxide. The carbon is oxidized to carbon mon-
oxide, and the carbon dioxide is reduced to carbon monoxide.

When carbon is burned in a limited supply of air carbon
monoxide is produced.

$$2C + O_2 = 2CO.$$
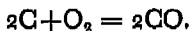
Expt. 140. Preparation of carbon monoxide. Fix up the
apparatus as shown in Fig. 96. The apparatus must be air-tight as
the gas is very poisonous. The experiment should also be performed
in a fume cupboard. A slow stream of carbon dioxide is passed

through water to remove the acid spray, and then over excess of char-coal heated to redness. The unchanged carbon dioxide is absorbed by the caustic soda in the wash-bottle. The gas is collected over water.

PROPERTIES OF CARBON MONOXIDE

The gas is colourless, odourless, and poisonous. It is only slightly soluble in water. When a lighted splint is put into a
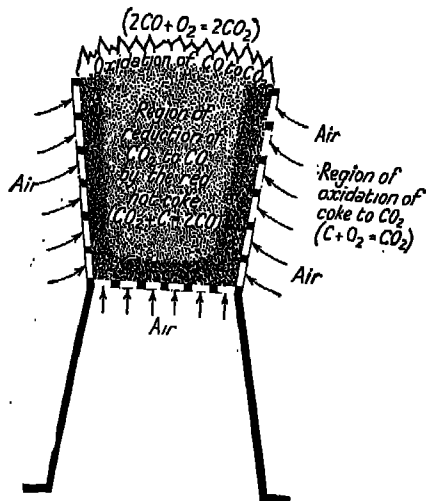
FIG. 97. A watchman's coke fire.

jar of the gas, the light is put out, but the gas burns with a blue flame. This flame is often seen in a clear red fire, especially in a coke fire. The air coming through the grating at the bottom of the fire (Fig. 97) comes into contact with red-hot coal or coke which burns to form carbon dioxide. This is drawn through the middle of the fire where it is reduced by the excess of red-hot coal to carbon monoxide. This gas, on reaching the top of the fire, again comes into contact with a plentiful supply of oxygen, and so burns to form carbon dioxide.

If carbon monoxide is mixed with half its volume of oxygen, and a light applied, it will explode. The product will of course

be carbon dioxide because the explosion is only rapid burning. Unlike carbon dioxide, carbon monoxide will not turn lime-water milky and is not absorbed by caustic soda or caustic potash. Carbon monoxide is a reducing agent. If ferric oxide is heated in a stream of the gas it is reduced to iron.

$$Fe_2O_3 + 3CO = 2Fe + 3CO_2.$$
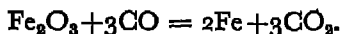
CARBON MONOXIDE AS A FUEL

Carbon monoxide is used extensively in industry as a fuel. It is usually mixed with other gases.

Producer-gas

If a limited amount of air is passed through hot coke, carbon monoxide is formed. In industry this operation is performed in a furnace called a generator or producer. The coke is burned, the air for this purpose entering at the bottom. The carbon monoxide produced, together with the nitrogen from the air, leaves the furnace by the flue and is carried to where it is needed as a fuel. There is an arrangement for charging the furnace with more coke without allowing any of the gas to escape. Producer-gas is a mixture of nitrogen and carbon monoxide, the nitrogen being about twice the volume of the carbon monoxide. It is clear that the heat to be obtained from burning producer-gas must be less than that which could be obtained by burning the coke which produced the gas. Some heat is wasted in the producer. This is partly compensated for by the fact that the gas is more easily transported to the places where it is wanted, than is coke.

Water-gas

In some gas-producer furnaces jets of steam are blown into the strongly heated coke. The carbon combines with the oxygen from the water to form carbon monoxide. Hydrogen is also formed.

$$H_2O + C = CO + H_2.$$

The mixture of gases thus formed is called water-gas. As both hydrogen and carbon monoxide burn, water-gas is an excellent

fuel. Steam cannot be passed through the furnace all the time, because the reaction causes the temperature to fall. At intervals the steam jets are stopped and air is passed in. The oxidation of the coke by the oxygen of the air raises the temperature. The steam jets are then opened again.

Water-gas is often added to coal-gas. This increases the percentage of hydrogen and carbon monoxide in the coal-gas and increases its value as a fuel. Unfortunately it makes it more poisonous, and of less value as an illuminant, although the latter is of little importance as a naked gas flame is now seldom used for illumination.

CARBON MONOXIDE AS A POISON

Carbon monoxide occurs wherever carbon or carbon compounds burn in a restricted supply of air. When petrol burns in the engine of a motor-car, some of the carbon in the petrol is oxidized to carbon monoxide which reaches the air by means of the exhaust pipe. If a motor-car engine runs in a closed space such as a garage, a dangerous amount of carbon monoxide enters the air. An escape of 'gas' in a building is dangerous because there is a large percentage of carbon monoxide in it. The fumes from any hot coke or other form of carbon should not be allowed to enter a closed room. Carbon monoxide may be present in the air in dangerous quantities without there being much indication of its presence. The gas is colourless and odourless, and the first symptoms of poisoning are not alarming. They are usually a slight headache followed by giddiness. Thus many accidents have been caused by the unsuspected presence of the gas. The gas is poisonous because it combines with the haemoglobin of the blood thus preventing oxygen from doing so. Respiration is therefore restricted and the patient dies from lack of oxygen although there is a plentiful supply available.

Anyone suspected of being poisoned by carbon monoxide should be kept warm and not allowed to walk about. Exertion and loss of heat use up the reduced supply of oxygen in the body very rapidly and every effort should be made to prevent this waste at such a critical time. If oxygen is available it

should be administered until the arrival of a doctor, who should be sent for immediately.

HYDROCARBONS

These are compounds of carbon with hydrogen. The simplest is methane, CH_4, often called marsh gas. Acetylene, C_2H_2, is well known as an illuminant. Paraffin wax, paraffin oil, and petrol are also well-known substances which are mixtures of hydrocarbons. They will all burn, and are used as fuels and for lighting purposes.

1. Marsh gas, methane, CH_4

If the bottom of a stagnant pond be stirred with a long stick bubbles of gas rise to the surface. These may be collected by holding in the water an inverted gas jar full of water. If a lighted taper be placed in the gas, the taper is extinguished, but the gas burns. The gas collected thus is not pure methane, and usually burns with a luminous flame. The pure gas burns with a blue non-luminous flame. The gas also occurs in coal-mines. It forms over 80 per cent. of the mixture of gases known as firedamp, which is present in most coal-mines. It is not poisonous, but when mixed with air it explodes if a light is applied. For this reason mines have to be efficiently ventilated to prevent any concentration of the gas, and all lights have to be guarded (p. 458). It is a constituent of coal-gas (about 30 to 45 per cent.), and forms over 40 per cent. of the so-called Natural Gas that comes from the petroleum springs of Russia and America.

When methane burns with a full supply of air, carbon dioxide and water are produced.

$$CH_4 + 2O_2 = CO_2 + 2H_2O.$$
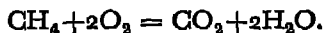
In a limited supply of air, some carbon monoxide is produced. This occurs in coal-mines, and sometimes causes the death of those who escape the effects of the explosion. The so-called After-damp is a mixture of nitrogen, carbon dioxide, and carbon monoxide.

2. Acetylene, C_2H_2

This gas is familiar to most people as an illuminant. It is still used for cycle lamps, but is now becoming less popular as the acetylene lamp is not so convenient as the electric lamp.

The gas may be prepared in the laboratory in the same way that it is generated in a cycle lamp.

Expt. 141. Preparation of acetylene. Put some pieces of calcium carbide, CaC_2, in a dry flask fitted with a tap funnel. This type
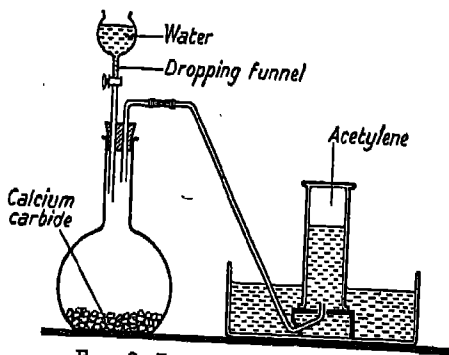
FIG. 98. Preparation of acetylene.

of funnel allows liquid to pass into the flask a little at a time without allowing any gas to escape (Fig. 98). Arrange for a delivery tube, and collect the gas over water. Do not allow any of the gas to escape, as it is poisonous. Its unpleasant smell, however, is some protection since we are aware of its presence. For this reason it is not so dangerous as carbon monoxide which has no smell.

$$CaC_2 + 2H_2O = C_2H_2 + Ca(OH)_2.$$
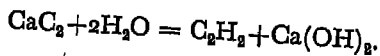
Prepared in this way the gas is colourless but has an unpleasant odour. It is said that when the gas is pure it has a pleasant sweet smell. It is not so dense as air and is slightly soluble in water, but not so soluble as to prevent its collection over water.

Expt. 142. Put a lighted splint into the gas. The gas burns with a smoky, luminous flame; but the light is put out. A light put into a

mixture of air and acetylene causes a violent explosion, so great care must be taken when using acetylene.

The difference between the flames of methane and acetylene is due to the fact that acetylene contains a higher proportion of carbon. The higher proportion of carbon in a hydrocarbon, the more luminous is the flame.

The flame in an acetylene lamp is not smoky, and a very bright light is given out. When acetylene is burned in the ordinary way the flame is smoky, because the whole of the carbon in the gas is not oxidized to carbon dioxide or carbon monoxide. Some of the carbon is unoxidized and comes off as free carbon. It is this black carbon that forms the smoke.

FIG. 99. An acetylene burner.

Examine the burner in an acetylene lamp (Fig. 99). Note the very fine hole through which the gas is forced. (There may be two of these holes arranged so that the jets of gas meet each other and produce a flat flame.) Note the holes through which air is drawn. This arrangement supplies the gas with enough air to oxidize the whole of the carbon. Thus a smoky flame is avoided. We shall understand this better when we have studied Flame (p. 200).

When acetylene burns in a plentiful supply of air or oxygen, the carbon is oxidized to carbon dioxide, and the hydrogen is oxidized to water.

$$2C_2H_2 + 5O_2 = 4CO_2 + 2H_2O.$$
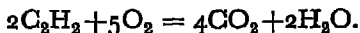
A very hot flame, called the Oxy-acetylene Flame, can be obtained by burning a mixture of oxygen and acetylene. This is done by means of a blowpipe, the mixture of gases issuing from the nozzle under pressure. The flame is hot enough to melt iron and steel and is used for fusing these metals. It is the hottest blowpipe flame known, hotter even than the oxy-hydrogen flame.

Acetylene is also formed when a bunsen burner 'strikes back' (p. 204).

·3. Petroleum

Crude petroleum oil occurs underground in many parts of the world. It is obtained by boring wells. When the bore hole reaches the oil underground, the oil often gushes up for days, but after a time it is necessary to pump the oil to the surface. The origin of petroleum is unknown. It has been suggested that it is the product of the decomposition of animal matter under conditions of heat and pressure.

Petroleum is a mixture of hydrocarbons, most of which are similar to methane. When petroleum is distilled many substances are obtained. The most volatile hydrocarbons are collected first, and then the less volatile, and so on, until towards the end, very heavy oils and solids are left. Petrol and paraffin are only two of the many important substances obtained from petroleum by distillation.

Paraffin oil is used as an illuminant and a fuel. When complete combustion takes place the paraffin is oxidized to carbon dioxide and water. In this respect it is like the other hydrocarbons. When used as an illuminant, the supply of air must be regulated so that complete combustion does not take place too quickly, because the light comes from the incandescent particles of carbon in the flame. This carbon is finally oxidized. If the supply of air is further restricted, the flame is smoky, all the carbon not being oxidized.

Paraffin wax is a solid form of paraffin. It consists of a mixture of hydrocarbons with higher melting-points than the hydrocarbons in paraffin oil. When a light is applied to the wick of a paraffin candle, the heat melts some of the wax, and the liquid travels up the wick towards the flame where it is vaporized and then oxidized (p. 206).

Petrol is also used as a fuel in stoves specially constructed for the purpose; but its chief use as a fuel is in the internal-combustion engine (p. 484). When completely oxidized the only products are carbon dioxide and water, but under conditions which exist in the internal combustion engine, carbon monoxide is also formed. Some of the carbon is not oxidized at all, and comes out free. The exhaust gases contain water

vapour. This is easily noticed when the engine is first started up and the cold exhaust tube condenses the water vapour. As the tube becomes warmer, the water vapour is not condensed.

COAL

Coal is familiar to us all. It is mined in many parts of the world, and England is fortunate in the possession of great coal-fields.

Coal is an impure form of carbon and is of vegetable origin. This is evident from an examination of the softer varieties of coal. In these, fossil plants can be recognized, and microscopic examination leaves no doubt that these varieties of coal have arisen from vegetable matter. Fossil plants are not found in the harder varieties, but it is believed that these too are of vegetable origin. All types of coal can be found showing a gradual transition from very soft to very hard varieties. The fossil plants and the vegetable structure gradually disappear as the varieties go from soft to hard.

When vegetable matter dies it is attacked by bacteria and the oxygen of the air and gradually decays. The gaseous products of the decomposition go into the atmosphere, and the solid parts eventually become part of the soil. If, however, decomposition occurs underground or in a swamp or bog where there is a limited supply of air, the whole of the carbon is not oxidized. This is believed to have occurred often in the past. Whole forests became submerged in bogs or swamps, and the wood decayed in a limited supply of air. In the course of ages these masses of vegetable matter sank into the earth, either gradually or by earthquake upheavals, and became subject to enormous pressure.

It is not surprising that we have many varieties of coal because the original deposits were different and the subsequent history of these deposits differed. The original deposits may have been trees, or soft vegetable matter. They may have been deposited in fresh water or sea water, or in bogs. They may have been subjected to enormous pressures for a long time, or to slight pressures for a comparatively short time.

Peat is supposed to be the first stage of the conversion of

vegetable matter into coal. There is 29 per cent. of carbon in peat. *Lignite* or *brown coal* is supposed to be the second stage. This variety has 43 per cent. of carbon. It burns with a smoky flame but does not give much heat. The third stage is *bituminous coal* with 63 per cent. carbon. This is harder than lignite, black and brittle. It burns with a yellow flame and gives out much more heat than lignite. The fourth stage is represented by *anthracite* with 86 per cent. carbon. It is hard and brittle, burns with difficulty, but with great heat.

Between each of the varieties mentioned there is a gradual transition. Peats can be found almost like lignite; and some bituminous coals are like lignite, while others are more like anthracite.

Besides carbon, the softer varieties contain water, and all varieties contain hydrocarbons. The carbon in the hydrocarbons is not included in the percentages given above.

COAL AS A FUEL

When coal is burned in an ordinary domestic grate, much of the carbon does not burn. This is evident in the amount of soot left in the chimney, and in the black smoke that escapes into the atmosphere. It is therefore a very wasteful way of burning coal.

Of the coal that does burn, the carbon is oxidized to carbon dioxide, though the presence of the blue flame in the fire indicates that carbon monoxide is formed at one stage of the oxidation (p. 182). The hydrogen is oxidized to water.

Much of the heat, with much of the unoxidized carbon and hydrocarbons, goes up the chimney. Only a fraction of the heat is used in warming the room. The coal furnaces used in factories are not so wasteful, but even now black smoke can be seen coming from factory chimneys. This is usually due to faulty stoking. Since the law regulating the emission of black smoke came into force, a better type of furnace has been generally introduced and more skilled stoking practised. This has resulted in a much more economical use of coal.

The emission of smoke from factory and domestic chimneys is the chief cause of the pollution of the atmosphere in many of

our towns and cities. We shall return to this subject when we have studied the chemistry of some of the substances causing the pollution.

THE DRY DISTILLATION OF COAL

Expt. 143. Arrange an apparatus as in Fig. 100. *A* is a hard glass boiling tube containing small pieces of coal (not powdered). The tube *A* slopes slightly downwards towards the mouth. It is fitted with a delivery tube *B* which leads almost to the bottom of another boiling
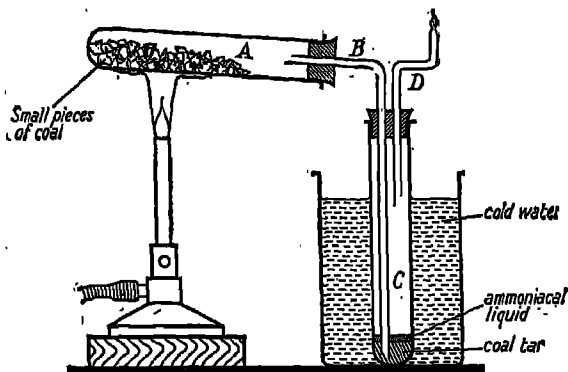
FIG. 100. The dry distillation of coal.

tube *C*. The tube *C* is immersed in a beaker of cold water. *D* is a tube leading from *C*.

Heat the coal in *A*, gently at first, but afterwards more strongly. The coal decomposes, and a heavy vapour appears in *A*, travels to *C*, some of it leaving at *D*. When the whole of the apparatus appears full of smoke, and there is a steady continuous stream issuing from *D*, apply a light at *D*.

The vapour burns with a yellow flame. Smell the air above the flame. Note the smell of burning sulphur. Blow out the light and smell the issuing gas. Note the smell of bad eggs. Hold a piece of filter paper moistened with lead acetate solution in the issuing gas. Note the black stain on the paper.

Re-light the vapour coming from *D* and allow it to burn until there is no more vapour coming from the coal. Allow the whole apparatus to cool and examine the contents. In *A* you will find some coke and probably some tar if you have not heated strongly. Note the two layers of liquid in *C*. The bottom layer is black and thick; the upper layer is more like water. Decant the upper layer into a test-tube.
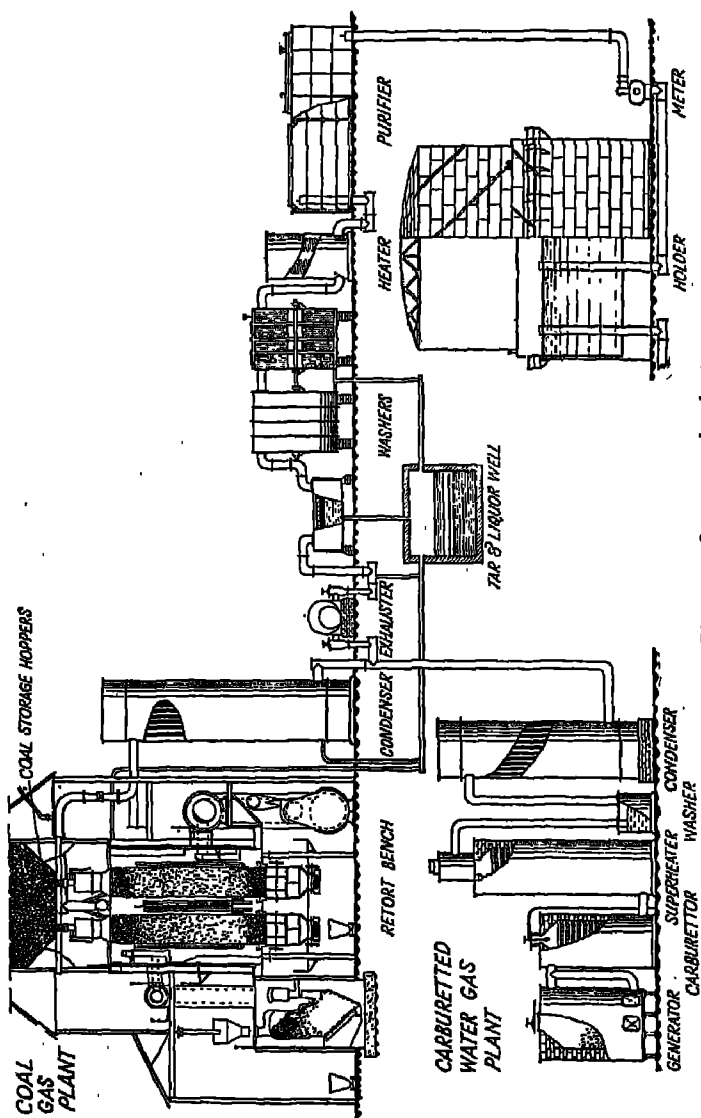
FIG. 101. Diagram of a gas-works plant.

Does it smell of ammonia? Add a little caustic soda solution and heat. Now smell the gas coming from the test-tube. Can you smell ammonia? (You should never smell an unknown gas except under instruction. Then you should smell carefully. Fill your lungs with air and then smell cautiously.)

The process of heating a substance dry so as to collect the volatile products of the decomposition, is called *dry distillation*, or *destructive distillation*. When coal is distilled in this way the products may be divided into three classes: solid residue, liquid residue, and gaseous products. The solid residue is coke which consists largely of free carbon. The gaseous product is called coal gas. The liquid residue consists of two parts: coal tar and the ammoniacal liquor. The dry distillation of coal is very important in modern industry and we shall study these different products separately.

COAL-GAS

Coal-gas is made on a large scale and distributed throughout the district in which it is made. The details of the method differ in different places, but the main outlines are the same.

Fig. 101 represents diagrammatically the plant used for the manufacture of the gas. Coal is heated in closed retorts. The volatile products travel up the ascension pipe which leads to the hydraulic main. This contains water and serves as a water seal, preventing the gas from flowing back into the retort. The gas passes through the liquid in the hydraulic main. In doing so it is cooled, and some tar and water are condensed. This tar is eventually led to the tar well. The gas then passes through long tubes called condensers. Here it is further cooled, and more tar is condensed and run into the tar well. The liquid here separates into two layers, the upper being the ammoniacal liquor or gas liquor, the lower being the coal-tar.

The gas which leaves the condensers contains impurities which must be removed. These are compounds of sulphur, carbon dioxide, and possibly some ammonia and tar. The gas we made in Experiment 143 was not freed from sulphur before it was burned, hence the smell of burning sulphur above the

flame. The smell of bad eggs was due to the presence of sulphuretted hydrogen (p. 212) in the gas.

The gas is drawn from the condensers by an exhaust pump and passed through 'scrubbers'. These are towers filled with pieces of stone and coke. From the top water trickles through the spaces between the coke and stones. The rising gas is broken up into bubbles and a large surface of the gas is exposed to water. Under these conditions the ammonia is dissolved out of the gas. Some, but not all, of the sulphur compounds are removed in the 'scrubbers'.

The gas then passes into the purifier, which contains slaked lime. This absorbs the sulphur compounds. The spent lime is sold as gas lime and is used in agriculture as a fertilizer. The gas then passes through further purifiers containing bog-ore (hydrated ferric oxide) which removes the last traces of sulphur compounds.

The purified gas is stored in gas holders which are usually huge covered cylinders floating in water. As the amount of gas stored increases, the gas holder rises in the water. The gas is distributed from the gas holders.

Coal-gas is a mixture, and its composition varies according to the kind of coal used and the method of distillation employed. Usually coal-gas contains about 50 per cent. hydrogen, methane 30 per cent., other hydrocarbons 5 per cent., and carbon monoxide 8 per cent. The remainder consists of nitrogen, oxygen, and carbon dioxide, which are considered impurities.

The advantages of coal-gas compared with coal as a fuel are that it is cleaner, it is much more easily transported, the temperature is more easily regulated; and it can be used more economically. When you have finished heating with coal-gas you can turn it off; with coal you must wait until the fire goes out, or draw the fire, both of which methods are wasteful.

SOME BY-PRODUCTS OF THE DRY DISTILLATION OF COAL

1. Coke

This is left in the retorts after the distillation is complete. It consists largely of the element carbon and some ash. It is used instead of coal for 'central heating'. A large amount is used by
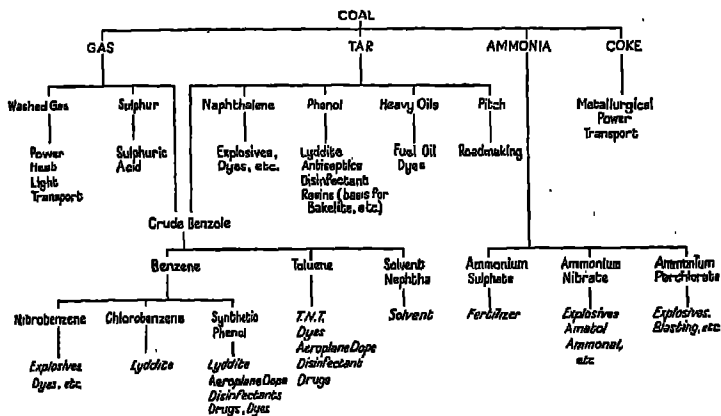
FIG. 102. A few of the products obtainable from the carbonization of coal. Nearly 200 substances have been isolated from tar up to the present time. The number of coal-tar derivatives obtainable is well over 2,000.

market gardeners for the heating of greenhouses. It does not burn so easily as coal, and therefore is not suitable for domestic purposes unless used in a special stove or grate providing a keen draught. It is used industrially as a furnace fuel and as a source of gaseous fuels:

(a) **Water-gas** is formed by passing steam over red-hot coke. When this gas is burned from a special burner, a short hot flame is obtained, used for welding and melting steel.

(b) **Producer-gas** or **fuel-gas** is obtained by passing air through beds of red-hot coke. This is used in factories and for 'gas engines'. It is also used in certain steel melting furnaces.

(c) **Semi-water-gas** is produced by combining the processes for making producer-gas and water-gas, that is, by

passing a mixture of air and steam over red-hot coke. It is used for the production of power by internal-combustion engines and for steam raising.

Coke is also used in the extraction of many metals such as iron and zinc from their ores.

There are two kinds of coke in general use. *Soft coke* is black and brittle, but burns with difficulty. It is used in blacksmiths' forges. *Hard coke* is dark grey and is not easily crushed. It is used in furnaces and in the extraction of metals from their ores.

2. Gas-carbon

This is a hard deposit which gradually accumulates on the sides of the retorts during the manufacture of coal-gas. It is almost pure carbon and is a good conductor of electricity, so it is used in the manufacture of electrical goods, e.g. the carbon rods of the electric arc, the carbon plates of electric batteries, and carbon brushes.

3. Ammonia

This occurs in the upper layer of liquid in the tar well. The liquid is boiled with milk of lime which expels the ammonia gas. This is passed through sulphuric acid to form ammonium sulphate.

$$2NH_3 + H_2SO_4 = (NH_4)_2SO_4.$$
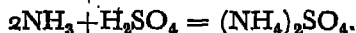
This substance is known commercially as 'sulphate of ammonia' and is a valuable fertilizer (Ch. XLVII).

4. Coal-tar or gas-tar

This is the black sticky substance found at the bottom of the tar well. It is used for protecting timber, and for water-proofing generally. Actually about two hundred different substances are present in coal-tar. A rough separation is made by 'fractional distillation', the receiver for the distillate being changed at certain temperatures. The tar is in this way separated into light oils, middle oils, and heavy oils.

The *light oil* distils over up to 170° C. It includes certain hydrocarbons such as benzene, toluene, and xylene (which are used as solvents) and is a source of aniline dyes, explosives, and

drugs. Benzene is also important as a fuel for motor engines.

The *middle oil* comes over between 170° C. and 230° C. This is the carbolic acid oil which is later further separated into naphthalene (the main constituent of 'moth balls'), and also
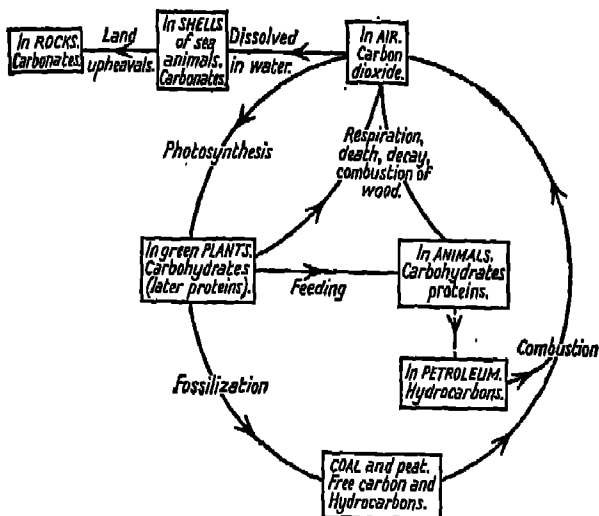
FIG. 103. The carbon cycle.

used as a starting-point for synthetic indigo, carbolic acid, and the cresols. The latter are used in the manufacture of many disinfectants and antiseptics, e.g. Jeyes's Fluid and Lysol. Picric acid, an important explosive, is manufactured from carbolic acid.

The *heavy oil* or creosote oil distils over at 230°–270° C. and is used to preserve timber.

The substance left in the retort when tar is distilled is called *pitch* and this is used for road making. *Asphalt* is a solution of pitch in a heavy tar oil.

THE MANUFACTURE OF PETROL FROM COAL

Coal contains carbon, and a much smaller percentage of hydrogen (among other things). Petrol is a mixture of hydrocarbons in which the amount of hydrogen present per given weight of carbon is approximately twice what it is in coal. It follows that if petrol is to be made from coal, the percentage of hydrogen must be increased. Hydrogen must also be supplied to combine with the unwanted oxygen of the coal so that it may be removed as water.

Powdered coal is mixed with a little alkaline iron oxide and bound with a suitable tar or oil. It is then heated with hydrogen to 400° C. under a pressure of 200 atmospheres. Under these conditions a mixture of liquid hydrocarbons is produced, and these are separated in the usual way by fractional distillation into petrol, heavier fuel oils, and lubricating oils.

The process is now carried on in England, but the petrol produced, though of high quality, cannot compete in the open market with 'natural' petrol, owing to the higher cost of production. It is hoped that some day the cost of production will be sufficiently low for it to be produced as cheaply as the 'natural' petrol. Meanwhile the production is encouraged by the government so that we could have an alternative supply, if for any reason the 'natural' petrol were unobtainable. During the Second World War the Germans increased their production of synthetic petrol enormously.

This process of manufacturing petrol from coal is known as the *hydrogenation of coal*.

QUESTIONS

1. Compare the physical properties of diamond, graphite, and lamp-black. What is the source of each and what are their uses? How can we prove that these are different forms of the same element carbon?

2. What are (a) hydrocarbons, (b) carbohydrates? How do they differ in composition? Give two examples of each.

3. How can carbon monoxide be prepared in the laboratory? What is (a) producer-gas, (b) water-gas? What are they used for in industry and how are they manufactured?

4. How does carbon monoxide act as a poison? State what first-aid treatment you would give to a person suffering from the effects of breathing the gas.

5. Explain the chemical actions which take place in (a) an internal-combustion engine using petrol, (b) a coke fire. Why should the exhaust gases from the engine and the fumes from the coke fire not be allowed to enter a closed garage or room?

6. What is (a) firedamp, (b) after-damp? Where do they occur? What happens when a hydrocarbon burns in (a) a plentiful supply, (b) a limited supply of air? Give equations.

7. What is petroleum? Where does it occur and how is it obtained? What is its supposed origin? Name some important substances obtained from petroleum.

8. Describe how coal was formed. What are the common varieties of coal and how do they differ from each other? Name two solid, one gaseous, and two liquid products of the dry distillation of coal.

9. What is the composition of coal-gas? What advantages has this fuel over coal?

10. During the dry distillation of coal many important by-products are obtained. Write a short account of some of these.

11. What volume of carbon dioxide is obtained when 100 c.c. of carbon monoxide is completely burnt in air? What volume of *air* is required for the complete combustion assuming all gases are measured at the same temperature and pressure?

XI

FLAME

THE PARTS OF A FLAME

In some previous experiments you have been cautioned that some liquids used were inflammable, and that the vapour of the liquid should not be allowed near a flame. It can be shown by experiment that flame is caused by the combustion of gases or vapours.

Expt. 144. Heat the end of a glass rod and bring it down slowly on to small quantities of various liquids such as carbon disulphide, alcohol, and benzene, placed in evaporating dishes. It will be noticed that a flame is produced when the rod is still some distance above the surface of the liquid. The temperature at which a substance takes fire is called the temperature of ignition of that substance.

We have seen that hydrogen burns in air or oxygen with a pale blue almost colourless flame. When closely examined it is found to consist of two parts (Fig. 104), an inner region and an outer region. The inner region is much cooler than the outer, and consists of unburnt hydrogen.

Expt. 145. Hold a piece of wire in various parts of the flame. It glows in the outer region but not in the inner. Place the end of a short piece of glass tubing in the inner region and light the gas issuing at the other end (Fig. 105).

We have previously proved that when hydrogen burns, water vapour is produced. We did this by allowing the flame to come in contact with a cold surface. This experiment shows that just outside the outer region water vapour is present, so it is evident that hydrogen must be combining with oxygen in the outer region.

Similarly with the flame of carbon monoxide, the inner region contains unburnt gas and the outer region consists of carbon monoxide combining with oxygen to form carbon dioxide.

THE BUNSEN BURNER

Expt. 146. Light a bunsen burner. Examine the different flames obtained by opening and closing the air-holes at the bottom of the tube. When the air-holes are open the flame gives out little light. In

Unburnt
hydrogen

FIG. 104.

Collar
Gas jet
Air hole
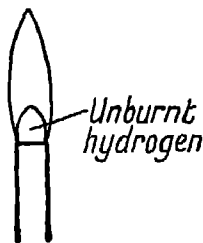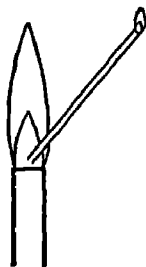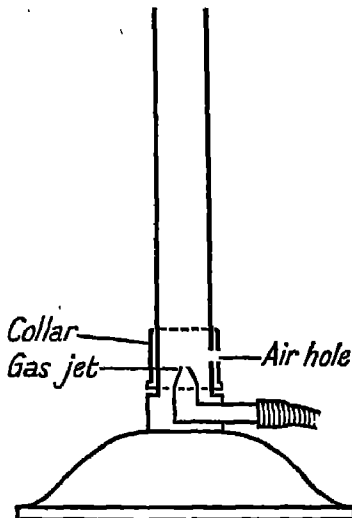
FIG. 105.   FIG. 106. The Bunsen burner.

strong sunlight this flame is scarcely visible. It is non-luminous. When the air-holes are closed the flame is yellow and luminous. Note the shape of the flame. Watch how the shape alters as the gas supply is altered by turning the gas tap.

Expt. 147. Take the bunsen burner to pieces by unscrewing the tube (Fig. 106). Note the small hole from which the gas issues. Light the gas at this nozzle. A thin pencil of luminous flame is obtained. Reassemble the burner and note the position of the air-holes opposite the point of entry of the gas. Note the arrangement for closing the air-holes.

To show that air enters at these holes and mixes with the gas, proceed as follows: Light the bunsen with the air-holes open. Put some iron filings on the bench near the bunsen and blow them into the air near the holes. Little sparks appear in the bunsen flame. The

iron filings enter with the air through the air-holes, travel up the tube
and are burnt in the flame, thus producing sparks. Sprinkle a few
iron filings on the bunsen flame. Sparks of the same kind are obtained.

THE BUNSEN FLAME
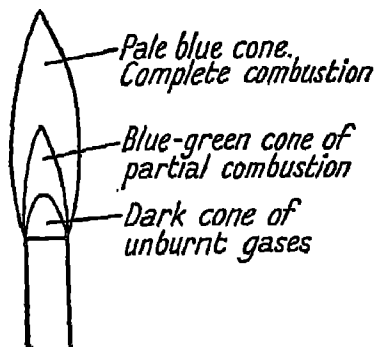
FIG. 107. The bunsen non-luminous flame.
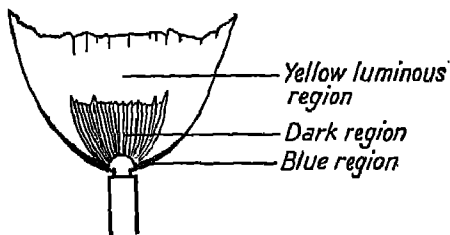
FIG. 108. A fish-tail burner.

Expt. 148. Light the bunsen and regulate the air-holes until the
bunsen 'roars'. Reduce the amount of air a little until the bunsen
just stops 'roaring'. Examine the flame and compare what you see
with Fig. 107.

The shape of the flame is roughly that of a cone, but the
shape depends on the pressure of the gas and the shape of
the burner. 'Fish-tail' burners (Fig. 108) are often used in
the laboratory and were once quite common in the home when
gas was the chief illuminant, and before the incandescent

mantle was in general use. The 'fish-tail' shape of the flame is caused by the shape of the jet from which the gas issues.
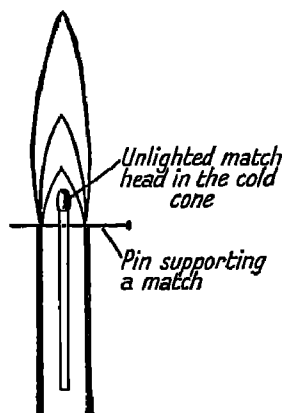
FIG. 109. An unlighted match in the cold region of a bunsen flame.

In the bunsen flame try to distinguish the following regions:
(1) a dark region in the centre,
(2) a blue-green region surrounding the dark region,
(3) a pale blue outer region.
What is happening in these regions?

Expt. 149. Hold a piece of stick for about a quarter of a minute horizontally, low down in the flame. Where is it burnt? Where is it unburnt? In what region of the flame was the unburnt part? Put another stick higher up in the flame and compare the results with the first. Is there any unburnt part? Push a pin through a match close to the head, and balance the pin on the top of the tube of the burner (Fig. 109). Now light the bunsen. The match head will not take fire if the flame is steady. The dark inner region is not so hot as the other regions.

Expt. 150. Hold a glass tube with one end in the dark region as shown in Fig. 105 and apply a light to the other end. Unburnt gas travels from the inner dark region and burns at the other end of the tube.

The dark region consists of unburnt gases (hydrogen, carbon monoxide, and hydrocarbons). In the blue-green region the

gases are partially oxidized. In the outer region the oxidation is completed to carbon dioxide and water. We have already shown that water is formed in the case of the hydrogen flame. Water is often seen on the outside of the vessel when a pan of cold water is placed on the flame of a gas stove. When a beaker or flask of cold water is placed over a bunsen flame the outside of the vessel clouds with water. This water comes from the flame.

The hottest part of the bunsen flame is at the edge of the outer region above the blue-green region. A small body held in the bunsen flame can be heated to a temperature above 1,000° C. The hottest flame for heating bodies held in it is that obtained by regulating the air so that the bunsen just does not 'roar'.

Expt. 151. Examine the burner of a gas stove or gas ring. Find out how the air is regulated. By altering the air supply different kinds of flames are possible. Adjust the air supply so that you get the best flame, that is one which is not 'roaring' but which is also non-luminous. Much gas is wasted by the inefficient use of the stove. It is worth a little trouble to see that the air supply is just right.

When a light is held to a bunsen burner or a gas ring and the gas then turned on, the flame often 'strikes back', and burns at the jet below the air-holes. An unpleasant smell is noticed. The bunsen has to be turned out and relit. The flame also strikes back when the air-holes are open and the gas is turned gently down. If you wish to turn a bunsen low you must also reduce the air supply. When a bunsen is burning normally with the air-holes open there is a tendency for the flame to travel down the tube, for it contains a mixture of gases which is combustible. Such a mixture in a vessel would explode when a light was applied. The mixture of gases, however, is travelling up the tube faster than the flame can travel down. Hence the flame remains steady. When the gas supply is reduced the speed at which the gases travel up the tube is reduced, and there comes a time when the speed of the gases up the tube is less than the speed of the flame down the tube. The flame then 'strikes back.'

THE LUMINOUS FLAME

We have seen that when the air-holes are closed the flame is luminous. Why does the flame lose its luminosity when air is allowed to enter at the air-holes? The explanation formerly given was as follows:

The flame emits light because of the presence of free carbon particles due to the decomposition of the hydrocarbons. When air is admitted complete oxidation of the carbon takes place and the flame loses its luminosity.

This cannot be the whole explanation, because it is found that the flame loses its luminosity when nitrogen is admitted through the air-holes. In this case there is no extra oxygen to oxidize the free carbon particles. It is thought that the cooling and diluting of the gases cause a loss of luminosity.

THE INCANDESCENT MANTLE

The old fish-tail luminous flame has now almost disappeared as a source of light in the home, and the incandescent mantle has taken its place where gas is used at all. The gas mantle is made by impregnating some fabric with thorium nitrate containing 1 per cent. of cerium nitrate. It is then heated and the nitrates are converted to the oxides. The mantle is now frail and cannot be carried about; it is therefore dipped in collodion. Before the mantle is first used the collodion is burnt away, and it will then emit a bright light when heated with the ordinary non-luminous gas flame.

Before putting a mantle in position the air should be carefully regulated so that an ordinary bunsen flame is obtained.

Examine an incandescent burner. Note the arrangement for the supply of air. Compare it with a bunsen burner.

THE CANDLE FLAME

Expt. 152. Examine the flame of a candle. Compare what you see with Fig. 110. Distinguish the following regions:
  (1) a dark region above the wick,
  (2) a large yellow luminous region above and at the sides of the. dark region,
  (3) a very faint narrow region surrounding the yellow region,

(4) a blue region below the wick.

Draw off some of the unburnt vapour from (1) with the aid of a narrow glass tube and light the vapour at the other end of the tube. (You must keep the tube hot otherwise the vapour will solidify.)

Hold a cold object, e.g. an evaporating basin containing cold water, in the flame and note the deposit of soot, showing the presence of particles of carbon in the flame.

Put some pieces of candle wax in a test-tube fitted with a cork and a piece of tubing drawn out to form a jet. Heat strongly with a
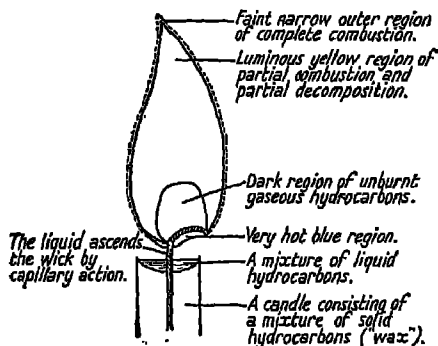
FIG. 110. The candle flame.

bunsen and when all the air in the tube has been driven out light the wax vapour coming out at the jet. Compare the flame with that of the candle. Care must be taken to see that the vapour does not solidify and block up the jet.

Blow out a candle and then immediately hold a flame near to, but not touching, the wick. The candle relights.

The 'wax', a mixture of solid hydrocarbons, is melted and the liquid travels up the wick by capillary action, in the same way that blotting paper soaks up water (Ch. XLVII). The liquid is then vaporized. The dark region consists of unburnt hydrocarbon vapours. In the yellow region the hydrocarbons are partly oxidized and partly decomposed into free carbon and hydrogen. The free carbon particles are hot and emit light. They are incandescent like the hot gas mantle. In the narrow outer region the carbon and gases are completely oxidized to carbon dioxide and water.

The blue region under the wick is very hot. Here the gases are completely oxidized. The heat of the flame causes an

upward current of air all round the flame, and the region under
the wick is thus well supplied with oxygen, thus enabling
complete oxidation to take place.

The wick is so plaited that it bends over, and its tip is thus
in the hot blue region. In this way the wick gradually burns
away. Formerly the wick was not plaited in this way and so
stood erect in the flame. As the wax of the candle gradually
burned away, the wick, which did not burn, became compara-
tively longer, and stood erect with its end in the yellow region.

Particles of carbon collected on
the end of the wick thus causing a
smoky flame and poor illumina-
tion. The wick had to be 'snuffed'
every few minutes. 'Snuffing' con-
sisted of removing the tip of the
wick and accumulated carbon by
means of a pair of 'snuffers' like
a pair of scissors. If the candle
was not snuffed the flame eventu-
ally went out.

BURNING OF A CANDLE

**Expt. 153. To find if there is
a change in weight when a candle
burns.** Push a piece of wire gauze
half-way down a cylindrical lamp
chimney (Fig. 111). On the gauze
place some small pieces of quicklime.
On top of this place a mixture of glass
wool and soda-lime reaching almost to
the top of the chimney. Fit to the other

FIG. 111.   To show that a
change in weight occurs when
a candle burns.

end of the chimney a cork through which many holes have been
bored. Place a candle on the cork in the position shown in the figure.
Suspend the lamp chimney from one arm of a balance and counter-
poise it. Remove the cork from the chimney, light the candle, and
replace the cork in position. Soon the balance shows an increase in
weight on the candle side.

The hydrocarbons forming the 'wax' of the candle are
oxidized to water and carbon dioxide. Oxygen of the air is
used for this. It enters through the holes bored in the cork.

The water formed is absorbed by the quicklime and so is prevented from falling back on the flame. The carbon dioxide is absorbed largely by the soda-lime but a proportion is doubtless absorbed by the lime after it has been slaked by the water. The glass wool is there merely to prevent the soda-lime from clogging the chimney. Unless a clear passage is left no air will enter through the holes in the cork and the candle will go out.
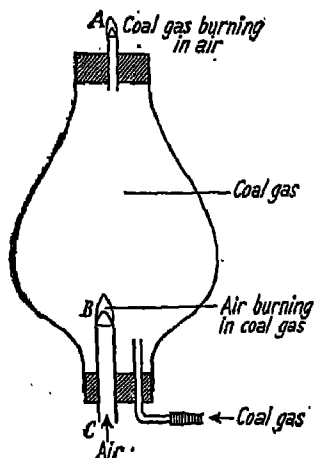
FIG. 112. Burning air in coal-gas.

The difference between the first weighing and the second is the weight of oxygen taken from the air by the candle in burning. Nothing is lost and oxygen is gained. The air must have lost exactly the amount of oxygen used in the combustion.

WHAT IS A FLAME?

In a flame we have seen that chemical actions between gases are going on. On the outside of the flame where the gases meet, new compounds are formed, and heat and usually light are produced. This seems to suggest that if we arranged to have air passing through a tube into an atmosphere of coal-gas, we could get a flame of air burning in coal-gas. This may be shown in the following experiment:

Expt. 154. Fix up a lamp chimney as shown in Fig. 112. The lower end is fitted with a cork carrying a wide glass tube BC and a narrower tube connected to the gas supply.

Turn on the gas supply and light it at A. Bring down over the flame at A a sheet of asbestos so as to reduce the flame, and while doing so apply a light to C. Remove the asbestos and the flame travels up from C to B.

We have, therefore, coal-gas burning in air at A, and air burning in coal-gas at B. If the end C of the wide tube is closed by the finger, air is prevented from entering, and the flame at B is extinguished, proving that here air was burning.

QUESTIONS

1. What is a flame? Illustrate your answer by reference to the bunsen and candle flames.

2. How would you adjust (*a*) a gas stove to give the hottest flame, (*b*) an incandescent gas light of the mantle type to give the best light?

3. How would you show that (*a*) water is one of the final products when coal-gas burns, (*b*) a candle increases in weight when it burns? Where does this extra weight come from?

4. Explain why the old practice of snuffing candles is no longer necessary.

5. Explain (*a*) why a bunsen burner sometimes 'strikes back', (*b*) why a loud 'pop' is sometimes heard a few seconds after the tap of a lighted gas stove burner is turned off.

6. Describe and explain what happens when the air-hole of a bunsen burner is gradually opened. Why should the air supply be reduced before a bunsen flame is lowered?

XII

SULPHUR AND PHOSPHORUS

THE ELEMENT SULPHUR

THIS element occurs free in many parts of the world, chiefly
in the neighbourhood of volcanoes. It is mixed with earthy
matter from which it is separated by heat. The sulphur melts
and runs from the earthy matter. It is afterwards purified by
distillation.

Sulphur also occurs free and almost pure in some parts of
America, chiefly in Texas and Louisiana. It occurs in thick
beds nearly 800 feet below the surface of the ground. It is
raised by·forcing water, super-heated to 160° C., down a bore-
hole. The sulphur is thus melted and forced to the surface
mixed with air and water. Most of the world's supply now
comes from America.

Sulphur also occurs in nature combined with other elements.

THE ACTION OF HEAT ON SULPHUR

Expt. 155. Powder some sulphur and heat it in a narrow test-tube.
(This may be considered as heating in the absence of air as very little
air will reach the mass of sulphur.) The sulphur melts. Note the
amber colour of the liquid and how it darkens in colour as the tempera-
ture increases until at the boiling-point it is nearly black. Note also
the change in the viscosity of the liquid. (When a liquid is thick like
treacle it is said to be *viscous*; when thin like water it is said to be
mobile.) It is mobile when first melted, then becomes very viscous,
and finally mobile again. Write an accurate description of the
changes in colour and viscosity as the temperature increases. Note also
what happens when the sulphur cools.

Expt. 156. Heat some more sulphur in a test-tube and when it has
reached the second mobile stage pour it into water. (You may wait
until the sulphur is boiling before pouring, but this is not necessary.)
When the sulphur you have poured into water has cooled sufficiently,
take it out and squeeze it between your fingers. Note that it can be
drawn out into transparent threads. This form of sulphur is known
as PLASTIC SULPHUR. As you work it in your fingers note that it
becomes harder, that is, less plastic. Leave it for a few days and
examine it again.

OTHER FORMS OF SULPHUR. ALLOTROPY

Expt. 157. Dissolve some powdered roll sulphur in carbon disulphide. (See p. 6 for the precautions to be used in using this liquid.) Filter through a *dry* filter paper and collect the filtrate in a watch glass. Allow the carbon disulphide to evaporate in a fume cupboard and use a lens to examine the crystals formed. This is known as RHOMBIC SULPHUR.

Expt. 158. Melt some sulphur in a fire-clay crucible. Allow it to cool until a thin crust has formed. Pierce two holes in the crust and pour out the still molten sulphur into water. Remove the crust and note the transparent needle-like crystals found on the sides of the crucible. These are crystals of MONOCLINIC SULPHUR or PRISMATIC SULPHUR.

It is obvious that the element sulphur can exist in more than one form, e.g. rhombic sulphur and monoclinic sulphur, which have different physical properties. When an element occurs in more than one form which differ in physical properties and sometimes in chemical properties, the different forms are said to be *allotropes* of the element and the phenomenon is known as *allotropy* (Greek *allos* = another, *tropos* = manner). Many elements have allotropes, the best known being carbon, sulphur, phosphorus, oxygen, and tin. We have already studied two allotropes of carbon, the diamond and graphite (p. 179), and we shall have to note the allotropes of phosphorus when we study that element (p. 221). The amorphous varieties of sulphur are not now considered as allotropes and amorphous carbon is supposed to consist of micro-crystals of graphite.

It is easy to show that allotropes of the same element consist of that element and nothing else, for if equal weights of the different allotropes are oxidized, then equal weights of the oxides are obtained. Further, one allotrope changes into another of the same element without change in weight.

The different properties of allotropes of the same element are accounted for by the fact that the molecules contain a different number of atoms in the molecule or a different arrangement in space of the atoms in the molecule. Thus, ozone, an allotrope of oxygen, has three atoms of oxygen in the molecule while ordinary oxygen has only two. The atoms

in a diamond molecule are arranged in a different pattern from those in graphite, and this arrangement accounts for the diamond being very hard and graphite being very soft.

USES OF SULPHUR

Sulphur is used for making sulphur dioxide and sulphuric acid, in the manufacture of matches and gunpowder, and for vulcanizing rubber. Raw rubber is sticky, but when heated with sulphur it hardens, the hardness increasing with the amount of sulphur added. Vulcanite and ebonite are prepared in this way. In medicine, sulphur is used in the treatment of skin diseases. Sulphur candles are sometimes burnt in rooms where there have been cases of infectious disease, to fumigate the room. Roses, vines, and hops are sometimes treated with powdered sulphur to destroy mildew.

FIG. 113. To show that hydrogen and sulphur combine.

SULPHURETTED HYDROGEN

Sulphur combines with hydrogen to form the compound Sulphuretted Hydrogen, or Hydrogen Sulphide, H_2S.

Expt. 159. To show that hydrogen and sulphur combine. Fit up the apparatus shown in Fig. 113. Heat the sulphur in the boiling tube until it boils, and then pass hydrogen from a generator into the sulphur. Put a piece of paper which has been dipped into lead acetate solution at the mouth of the delivery tube. The lead acetate turns black. There is a smell of bad eggs.

The gas that turns lead acetate paper black in Experiment 159 is hydrogen sulphide. The hydrogen combines with sulphur to form hydrogen sulphide. We say that the compound

has been synthesized. Not much hydrogen sulphide can be formed in this way. The greater part of the gases coming out of the mouth of the delivery tube consist of uncombined hydrogen and sulphur vapour. The gas is colourless and its smell is that which we associate with the smell of bad eggs. This is because eggs contain a sulphur compound, and when this decomposes, hydrogen sulphide is formed.

If a quantity of the gas is required, it may be prepared by the action of dilute hydrochloric acid on ferrous sulphide and collected by downward delivery or over hot water.

$$FeS + 2HCl = FeCl_2 + H_2S.$$
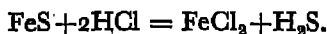
In Experiment 50 we heated a mixture of iron and sulphur and found that the product could not be separated into its original constituents. The product did not give off hydrogen when treated with dilute hydrochloric acid, but gave sulphuretted hydrogen. The solid formed by heating iron and sulphur together is, in fact, *ferrous sulphide*.

FIG. 114. Kipp's apparatus for the production of sulphuretted hydrogen.

Sulphuretted hydrogen is often used in the laboratory and it is convenient to have a ready supply. This is effected by the use of Kipp's Apparatus (Fig. 114). When the tap A is opened the pressure in B is reduced, and dilute hydrochloric acid rises into B where it comes in contact with the ferrous sulphide. Hydrogen sulphide is produced and passes out through the tap A. When no more gas is required the tap is turned off, and the gas produced in B after this cannot escape. The pressure in B is thus increased and the acid forced into C and up the tube into D. On turning the tap on again hydrogen sulphide is at once available and the same process is repeated.

Kipp's Apparatus can of course be used for supplying some other gases, e.g. carbon dioxide and hydrogen.

PROPERTIES OF HYDROGEN SULPHIDE

Hydrogen sulphide is easily soluble in cold water and so must be collected by downward delivery or over hot water. The gas is slightly denser than air.

Expt. 160. Put a light into a jar of the gas. The latter burns with a blue flame and there is a deposit of sulphur on the inside of the jar.

When the gas burns in a plentiful supply of air, e.g. at a jet, both the hydrogen and the sulphur are oxidized, the former to water and the latter to sulphur dioxide:

$$2H_2S + 3O_2 = 2H_2O + 2SO_2.$$
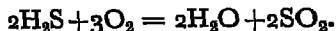
When the air is limited, as in our experiment, the hydrogen is oxidized, but some of the sulphur is not, and is deposited as free sulphur.

$$2H_2S + O_2 = 2H_2O + 2S.$$
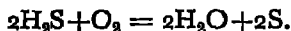
Expt. 161. Put some water into a jar of the gas and shake. Add blue litmus solution. It turns red.

The gas is an acid because its hydrogen can be replaced by a metal. The salts of sulphuretted hydrogen are called *sulphides*. Some sulphides may be made by synthesis from the elements; others are made by passing hydrogen sulphide through a solution of a salt, e.g. when hydrogen sulphide is passed through a solution of lead nitrate, a black insoluble substance, lead sulphide, is obtained.

$$Pb(NO_3)_2 + H_2S = PbS + 2HNO_3.$$

It is for this reason that lead acetate paper turns black in the presence of hydrogen sulphide, lead sulphide being formed. The lead sulphide may be separated from the dilute nitric acid by filtering.

SULPHUR DIOXIDE, SO₂

This gas is always formed when sulphur, and some compounds of sulphur, burns in the air; but if we wish to collect a quantity of the gas, it is better to prepare it by the action of copper on sulphuric acid, which is a compound of sulphur.

Expt. 162. Preparation of sulphur dioxide. Fit up the apparatus as in Fig. 115. Cover the copper turnings with concentrated sulphuric acid. (*Warning*. Before using concentrated sulphuric acid, read p. 219, for the precautions to be taken. Concentrated sulphuric acid is dangerous if carelessly handled.) Heat gently. Collect the gas evolved by downward delivery.

When the jar is thought to be full, put a lighted splint in its mouth. If the light is put out, the jar is full of the gas.

Besides sulphur dioxide, copper sulphate and water are formed, according to the equation:

$$Cu + 2H_2SO_4 =$$
$$CuSO_4 + 2H_2O + SO_2.$$
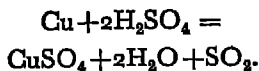
The contents of the flask appear black. The copper sulphate you used before was blue. Why then are not the contents of the flask blue? It depends on a property of sulphuric acid which we shall

FIG. 115. Preparation of sulphur dioxide.

study in detail later. It is very ready to combine with water. Copper sulphate is blue in crystalline form; that is, when combined with water of crystallization (p. 224), but is white when deprived of this water. The concentrated sulphuric acid unites with the available water, so the sediment in the flask should be white. However, a small quantity of a black solid, cuprous sulphide, is formed in addition. If the contents of the flask are poured into a large quantity of water a blue liquid is obtained.

The sulphur dioxide collected is colourless, fumes slightly in moist air, and has a smell of burning sulphur. It is heavier than air.

Expt. 163. Put a lighted splint into a jar of the gas. The gas does not burn and the light is put out. Pour a little water into a jar of the gas and shake. Test the liquid with litmus solution. It turns red.

The gas is soluble in water and the solution is acid.

$$SO_2 + H_2O = H_2SO_3 \text{ (Sulphurous acid)}.$$
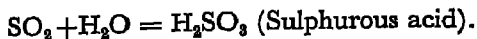
This acid is *unstable*, that is, easily decomposed. If a solution of sulphurous acid is heated, it decomposes into water and sulphur dioxide.

The gas is very easily liquefied by being passed through a vessel immersed in a freezing mixture. It can be liquefied at the ordinary temperature under pressure, and is sold in siphons similar to those in which soda water is sold. The gas is used as a disinfectant, and this way of carrying it to the affected place is very convenient. On turning the screw stopper, sulphur dioxide gas issues from the siphon. If liquid sulphur dioxide is required, the siphon is turned upside down and the stopper released in this position. When it is desired to disinfect a room all the windows and doors are closed and liquid sulphur dioxide poured on to dusters and left to evaporate, which it does very quickly. Sulphur candles are often burned instead. This produces sulphur dioxide and has the same effect.

Expt. 164. Put some rose petals in a jar of the moist gas. The petals soon lose their colour.

They have been bleached by the sulphurous acid. The gas, in the presence of water, is used commercially as a bleaching agent for straw, wool, and silk.

SULPHUR TRIOXIDE, SO_3

Sulphur dioxide will not burn in oxygen to form sulphur trioxide, but if a mixture of sulphur dioxide and oxygen is passed over heated platinized asbestos, thick white fumes of sulphur trioxide are formed, and may be collected by passing them into a flask immersed in a freezing mixture. Platinized asbestos is asbestos containing between its fibres finely divided platinum. The latter acts as a catalyst.

$$2SO_2 + O_2 = 2SO_3.$$
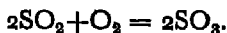
Sulphur trioxide combines readily with water to form sulphuric acid.

$$SO_3 + H_2O = H_2SO_4.$$
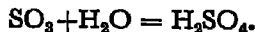
Sulphur trioxide, like sulphur dioxide, is an acidic oxide. This fact is now used in the manufacture of sulphuric acid.

SULPHURIC ACID

This is one of the most important acids and is used in enormous quantities in the chemical industry. It is used in the manufacture of hydrochloric acid and nitric acid. When added to three or four times its volume of water it is used in
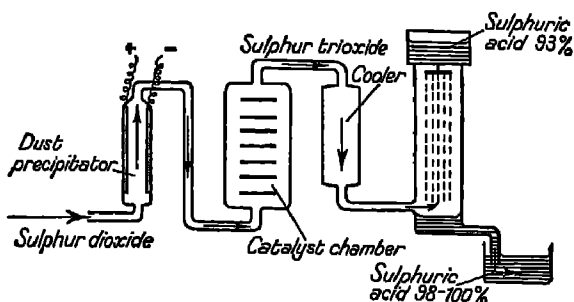
FIG. 116. The Contact Process.

accumulators. It is also used in the manufacture of the fertilizers ammonium sulphate and 'superphosphate' (a phosphate of calcium); for the manufacture of many dyes and explosives; in petroleum refining, and in the manufacture of celluloid for films.

Sulphuric acid was formerly made by heating ferrous sulphate, commonly called green vitriol, hence the common name for sulphuric acid, oil of vitriol. Vitreous means glassy, and green vitriol looks like green glass. Sulphuric acid, or oil of vitriol, is a heavy, colourless, oily liquid. There are now two common methods of manufacture, the Contact Process and the Lead Chamber Process.

In the CONTACT PROCESS the platinized asbestos method is used (Fig. 116). The sulphur dioxide and air, first cleansed from dust, are passed into chambers where they come in contact with the heated platinized asbestos. The sulphur trioxide formed is passed into concentrated sulphuric acid into

which water trickles at such a rate that the concentration is kept the same.

It is necessary to cleanse the gases from dust because it is found that the catalyst will not work long if contaminated, especially by substances such as arsenic. It is not clear why this is so, but experience shows that it is true. It is said that the catalyst is 'poisoned', but this is merely a picturesque way of describing it.

It is also found that when sulphur trioxide is passed into water much of it escapes without dissolving. It is to prevent this that the sulphur trioxide is dissolved into the acid, and the necessary amount of water added at the same time.

The LEAD CHAMBER PROCESS, in principle, is the same as the contact process but its details are widely different. As in the contact process, sulphur dioxide is made to combine with oxygen by the aid of a catalyst, and the resulting sulphur trioxide combines with water to form sulphuric acid. The catalyst employed in this process is the gas nitric oxide, NO. This substance combines easily with the oxygen of the air to form nitrogen peroxide, NO_2.

$$2NO + O_2 = 2NO_2.$$
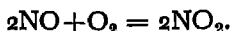
Sulphur dioxide, steam, air, and some nitric oxide meet in large lead-lined chambers (Fig. 117). The nitric oxide combines with the oxygen of the air to form nitrogen peroxide. This then gives up the oxygen to the sulphur dioxide which thus becomes sulphur trioxide, while the nitrogen peroxide again becomes nitric oxide.

$$NO_2 + SO_2 = NO + SO_3.$$
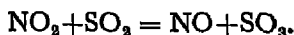
As the sulphur trioxide is formed, it combines with the water present in the form of steam and becomes sulphuric acid.

Such, at least, is one explanation of the process, but others have been given. The above explanation must be looked upon only as the simplest of many possible explanations.

Sulphuric acid made by this process has a concentration of about 70 per cent. If a greater concentration is needed, the acid is boiled in vessels made of a material which is not

attacked by the acid. Much of the acid is sold without further
concentration as it is quite suitable for many processes at this
strength.

The contact process is slowly supplanting the older chamber
process, but the latter can still hold its own especially when a
concentrated acid is not required.
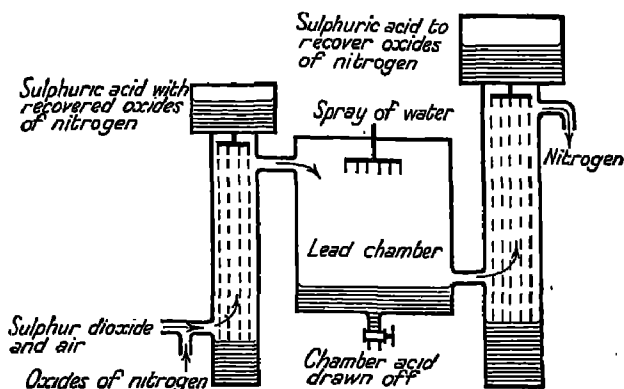
FIG. 117. The Lead Chamber Process.

PROPERTIES OF SULPHURIC ACID

When concentrated sulphuric acid is mixed with water
there is a considerable rise in temperature. Sometimes the
temperature rises so quickly that steam is produced which, in
escaping, spurts the acid out of the vessel. As sulphuric acid is
very corrosive and injurious to the skin, it must be mixed with
water with great care. Water should never be added to the
acid, but acid added in small quantities to the water, stirring
all the time.

Sulphuric acid is hygroscopic. For this reason it is sometimes
used for keeping the air in a desiccator dry. It is also used for
drying some gases by passing them through the acid. Of course,
it must not be used for gases which act chemically on the acid,

e.g. it cannot be used for drying ammonia as this gas combines with sulphuric acid.

Sulphuric acid will not only absorb moisture readily, it will take water from some compounds which contain hydrogen and oxygen in the same proportion as they exist in water. This is sometimes called 'removing from a compound the elements of water'.

Expt. 165. Add a little concentrated sulphuric acid to a little of the following substances, each in a separate test-tube: wood, paper, sugar, and starch. Note that each chars.

This is due to the carbon left after the elements of water have been removed. Each of these substances is either a compound, or contains compounds, consisting of carbon, hydrogen, and oxygen, the last two being in the same proportion as in water.

If concentrated sulphuric acid gets on the skin and is not quickly washed off, the skin is damaged and a scar is left which is likely to be permanent. The acid probably acts on the skin as it does on sugar and starch, taking from it the elements of water and thus destroying it as skin. Never rub acid off the skin; always wash it off with plenty of water.

On p. 148 we learned that sulphuric acid contained hydrogen which could be replaced by a metal to form salts. This is true of all acids. The salts formed from sulphuric acid are called *sulphates*. Again, like all other acids, sulphuric acid will turn blue litmus solution red. Some important sulphates are given on p. 227.

SULPHUR COMPOUNDS IN THE ATMOSPHERE

A sulphide of iron, iron pyrites, FeS_2, often occurs in coal. It looks like streaks of gold among the coal. (It is sometimes called 'Fool's gold'.) When the coal is burned, the iron pyrites is oxidized to sulphur dioxide and ferric oxide, Fe_2O_3.

$$4FeS_2 + 11O_2 = 2Fe_2O_3 + 8SO_2.$$
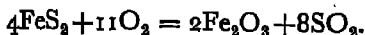
The ferric oxide remains among the ashes as reddish scales. The sulphur dioxide gas goes up the chimney and escapes into the atmosphere. If a shovel of burning coal is carried through a room, the smell of sulphur dioxide is noticeable. The air in

stokeholds and furnace rooms nearly always smells of sulphur dioxide.

The air of towns, especially industrial towns, contains traces of sulphuretted hydrogen as well as sulphur dioxide. Silver ware very soon tarnishes in the air of these towns, because silver sulphide, which is black, appears on the surface of the silver. Buildings painted with 'white lead' paint also turn black, because the lead in the lead carbonate, which is the basis of white lead paint, is changed to lead sulphide, which is black. This also accounts for old paintings turning a darker colour. It is not 'age' which does this, but the sulphur compounds in the air. Most of the pigments contain white lead, and the lead sulphide formed darkens the pigment. The 'restoration' of such paintings consists partly in changing the black lead sulphide into a white compound of lead.

Sulphur also occurs in india-rubber, and a silver coin kept in contact with a piece of india-rubber for a few days turns black. The sulphur in an egg or in mustard will cause a silver spoon to tarnish. There are also cases of tarnishing due to direct combination and not to sulphuretted hydrogen and sulphur dioxide.

PHOSPHORUS

Phosphorus is important because, for one reason, it is an essential constituent of all living matter. Unlike sulphur, it does not occur free in nature, but it occurs as calcium phosphate in minerals. Bones contain about 70 per cent. of calcium phosphate. As all plants take it from the soil, the farmer has to put it back if the fertility of the soil is to be maintained (Ch. XLVII). The sources of phosphorus are not unlimited and one of the problems of agriculture in the future will be to conserve the phosphorus compounds in the soil.

The element phosphorus was formerly manufactured from bone ash, and is still to some extent, but most phosphorus is now manufactured from the mineral phosphates by heating them with sand and coke in an electric furnace.

There are two common varieties of phosphorus, *white*

phosphorus and *red phosphorus*. (White phosphorus is commonly called yellow as most laboratory specimens are slightly yellow.) It is a waxy crystalline solid which takes fire in air at about 45° C. It should never be touched with the fingers as their heat might cause it to take fire. It is very poisonous, and so are the fumes of phosphorus pentoxide which are formed when it burns in air. It is insoluble in water but soluble in carbon disulphide. It is kept under water.

Red phosphorus appears amorphous but is probably crystalline. It is insoluble in carbon disulphide and is non-poisonous. It is less active than white phosphorus and for this reason is more commonly used in the laboratory.

Matches originally contained white (yellow) phosphorus, but as the workmen engaged in their manufacture suffered from the poisonous fumes, causing decay of the jaw, its use for this purpose has been prohibited. *Safety matches* contain a mixture of antimony sulphide, gum, and an oxidizing agent (usually potassium chlorate) to supply the oxygen for combustion, and red phosphorus is on the side of the box where the matches are struck. 'Strike anywhere' matches contain phosphorus sulphide, gum, and an oxidizing agent. The wood splints are dipped in boric acid solution to prevent the used matches from continued glowing.

There are several allotropes of phosphorus (p. 211). White phosphorus is the commonest. Formerly red phosphorus was considered an allotrope, but recent workers do not consider it so. The subject of allotropy is receiving much attention at present and is becoming too complicated for elementary treatment.

QUESTIONS

1. How and where does sulphur occur free in nature? How is it extracted? What are the uses of the element?

2. How would you show that sulphuretted hydrogen is a compound of sulphur and hydrogen and nothing else? How is the gas usually prepared in the laboratory and what are its uses?

3. Compare the properties of sulphur dioxide and sulphuretted hydrogen. In what form can sulphur dioxide be bought? What are its uses?

4. Give, without manufacturing detail, the two methods of the industrial preparation of sulphuric acid, and state the advantages of both. What part do catalysts play in the manufacture?

5. Give an account of the properties of concentrated sulphuric acid. What are its uses?

6. How do sulphur compounds get into the air and what are the effects of this pollution? Account for the smell of burning sulphur often noticed in a furnace room.

7. What are the sources of phosphorus and why do farmers occasionally put phosphates into the soil? What part does phosphorus play in the manufacture of matches?

8. What do you understand by the term allotrope? Illustrate your answer by reference to the elements carbon, sulphur, phosphorus, and oxygen.

XIII

ACIDS, BASES, AND SALTS

WE have already learned something about some acids and their salts, e.g. sulphuric acid and sulphates, carbonic acid and carbonates, and sulphuretted hydrogen and sulphides. We have also learned the relation between an acid and its salts, but so far we have said very little about alkalis, and nothing at all about a very important class of substances called bases. Let us first learn some more about salts.

SODIUM CARBONATE

This is a salt of carbonic acid. In the form of washing soda it is very common. Pure sodium carbonate is a white solid. If dissolved in water and then allowed to crystallize, colourless crystals of washing soda are formed.

Expt. 166. To observe the effects of heat on certain crystal-line salts. Heat crystals of washing soda, alum, dry common salt, cobalt chloride. If any liquid is formed test it with white copper sulphate to see whether it is water or not (p. 152). Note the state of the solid left in the tube.

It will be found that some crystals give off water when heated, and turn into a powder, sometimes with change of colour. On the other hand some crystalline salts, e.g. common salt, do not give off water when heated.

Such water is known as *water of crystallization* because the crystalline form of the substance depends on the presence of this water. The water is actually *combined* with the substance in a perfectly definite proportion and is not just mixed. A definite chemical action takes place between the water and the substance, as is shown by the evolution of heat when water is added to white copper sulphate, indicating that a chemical change has taken place.

Substances containing water of crystallization are often known as *hydrated substances*. The dry salt, after removing water of crystallization, is known as the *anhydrous* salt. It must be remembered that there are other crystalline substances

which are not salts, such as iodine and sulphur. These are elements and do not have water of crystallization.

Washing soda consists of sodium carbonate combined with water. The formula of the molecule is $Na_2CO_3.10H_2O$. Washing soda is hydrated sodium carbonate. When strongly heated it becomes anhydrous sodium carbonate, Na_2CO_3.

No doubt you have often noticed that washing soda in the kitchen has a floury appearance. If you wash this floury substance off the crystals and leave them exposed to the air, the floury substance appears again. This powdery substance is another hydrated form of sodium carbonate, the formula of which is $Na_2CO_3.H_2O$. The washing soda has *effloresced*. Each molecule of sodium carbonate has lost nine molecules of water of crystallization, but has retained one. The floury powder is called the mono-hydrate, and the washing soda is called the deca-hydrate. Only some salts having water of crystallization effloresce on exposure to the air.

Sodium carbonate is derived from carbonic acid, H_2CO_3, by the replacement of the hydrogen of the acid by the metal sodium. Each atom of hydrogen is replaced by one atom of sodium, so two atoms of sodium are needed to make one molecule of sodium carbonate. This salt is not made by the direct replacement of the hydrogen by the sodium, but in more indirect ways.

Expt. 167. Pass carbon dioxide for a long time through a solution of caustic soda (sodium hydroxide, NaOH). Concentrate the liquid by evaporation and allow the crystals to form as the liquid cools. These are crystals of sodium carbonate (washing soda).

$$2NaOH + CO_2 = Na_2CO_3 + H_2O.$$
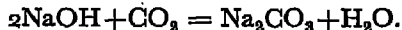
A solution of sodium carbonate in water turns red litmus solution blue, but sodium carbonate is not an alkali but a salt. We shall see later what an alkali is; meanwhile we may note that some salts dissolved in water have an alkaline reaction with litmus.

When an acid is added to sodium carbonate carbon dioxide is evolved and a salt of the acid produced.

$$Na_2CO_3 + H_2SO_4 = Na_2SO_4 + H_2O + CO_2.$$
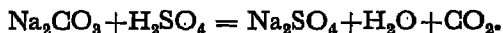
Here the hydrogen of the sulphuric acid replaces the sodium of the sodium carbonate and forms carbonic acid. This acid, however, is unstable and decomposes into water and carbon dioxide.

$$H_2CO_3 = H_2O + CO_2.$$
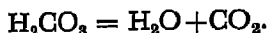
Large quantities of washing soda are used for domestic purposes. It emulsifies oil and grease (p. 276), and in this way oily and greasy objects are cleaned much more easily.

Sodium carbonate is also used for water softening (p. 278), glass making, paper making, and in the manufacture of many chemicals.

NORMAL SALTS AND ACID SALTS

Expt. 168. Pass carbon dioxide for a long time through a saturated solution of sodium carbonate in which there is present some solid sodium carbonate. The solution becomes milky, owing to the formation of a finely divided solid, known as bicarbonate of soda.

This substance is sodium bicarbonate, or sodium hydrogen carbonate, or bicarbonate of soda, or baking soda. Its formula is $NaHCO_3$. This salt, like sodium carbonate, is also derived from carbonic acid; but in this case only part of the replaceable hydrogen has been replaced. There is still one atom of replaceable hydrogen per molecule not replaced. In other words this substance still contains hydrogen which can be replaced by a metal. It has therefore the character of an acid. At the same time it is undoubtedly a salt because the sodium has replaced hydrogen from the acid. Such a substance is called an *acid salt*. A *normal salt* is one in which all the replaceable hydrogen of an acid has been replaced by a metal; an acid salt is one where only part of the replaceable hydrogen has been replaced.

Sodium sulphate, Na_2SO_4, is a normal salt of sulphuric acid; sodium bisulphate, $NaHSO_4$, is an acid salt of that acid. Acids like carbonic acid and sulphuric acid are said to be *dibasic* because they have two atoms of replaceable hydrogen per molecule; an acid like hydrochloric acid, HCl, is *monobasic* having only one atom of replaceable hydrogen per molecule.

Calcium carbonate is a normal salt of carbonic acid. One

atom of the element calcium is able to replace two atoms of hydrogen. Hence the formula for calcium carbonate is $CaCO_3$. We have seen, however, that calcium bicarbonate is formed when calcium carbonate is acted on by water and carbon dioxide. What then will be the formula of calcium bicarbonate? If only part of the hydrogen is to be replaced the part of the molecule left must be—HCO_3. As one atom of calcium can replace two atoms of hydrogen there will have to be two —HCO_3 groups in the molecule of calcium bicarbonate, $Ca(HCO_3)_2$.

SOME IMPORTANT SALTS OF SULPHURIC ACID

Sodium sulphate, $Na_2SO_4.10H_2O$, is used in medicine as Glauber's Salt.

Magnesium sulphate, $MgSO_4.7H_2O$, is used in medicine as Epsom Salts.

Ammonium sulphate, $(NH_4)_2SO_4$, obtained chiefly as a by-product of the manufacture of coal-gas, is used extensively as an artificial manure (Ch. XLVII).

Calcium sulphate, $CaSO_4.2H_2O$, is found in nature as gypsum. When heated to about 120° C. it loses some of its water of crystallization and is then called Plaster of Paris, $(CaSO_4)_2.H_2O$. Water is mixed with this to form a paste which quickly 'sets' and once more becomes $CaSO_4.2H_2O$. In setting it also expands a little. It is used by dentists for making plaster casts and also by doctors for embedding limbs.

Copper sulphate, $CuSO_4.5H_2O$ (blue vitriol), is very poisonous to some of the smaller forms of life. Mixed with sodium chloride it is used for killing 'blight' on certain crops. It is also used for copper plating.

BASES

We have seen (p. 146) that when hydrochloric acid acts on a metal hydrogen is evolved and a chloride of the metal is formed. Taking magnesium as an example:

$$Mg + 2HCl = MgCl_2 + H_2.$$

Let us now see what happens when hydrochloric acid acts

on a metallic oxide. We shall take magnesium oxide as an example.

Expt. 169. To a little hydrochloric acid in a test-tube add a pinch of magnesium oxide. It dissolves. Add more until no more dissolves on warming. Filter to separate the solution from the undissolved magnesium oxide. Evaporate the filtrate to dryness. A white solid is obtained which is magnesium chloride, the same substance that is obtained when the metal magnesium is dissolved in the acid. Note that no hydrogen is evolved.

The hydrogen that is displaced from the acid combines with the oxygen of the oxide to form water. This is not obvious in the experiment because there is water present all the time and it is impossible to see if any other water is formed. To do this we should have to pass dry hydrogen chloride over heated magnesium oxide. Then we should see the water formed. The equation for the action is:

$$MgO + 2HCl = MgCl_2 + H_2O.$$
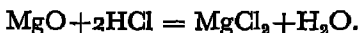
Compare this equation with the one above and you will see the connexion between the two.

When a substance acts on an acid to give a salt and water and nothing else, that substance is called a *base*. The relation between the three classes of substances is:

$$Base + Acid = Salt + Water.$$

Expt. 170. Repeat the above experiment using calcium oxide. The salt obtained in this case is calcium chloride. (You will not get a nice dry powder on evaporation but a sticky substance.)

$$CaO + 2HCl = CaCl_2 + H_2O.$$

If calcium hydroxide is used instead, the products of the reaction would still be calcium chloride and water only.

$$Ca(OH)_2 + 2HCl = CaCl_2 + 2H_2O.$$
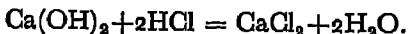
Calcium hydroxide is also a base, because with an acid it gives a salt and water only.

It is clear from the above that it is possible for only two classes of substances to be bases—oxides of metals and hydroxides of metals. We have seen that calcium hydroxide is made by

the action of calcium oxide on water. Not all hydroxides can be made directly from the oxides in this way, but many of them can. No matter how an hydroxide is made, chemically it may be looked upon as the oxide combined with water. For this reason, when an acid acts on an hydroxide, we expect to get the same products as when the acid acts on the corresponding oxide.

All oxides of metals are not necessarily bases. Some oxides, e.g. lead peroxide, give with an acid a salt, water, and something else. They are not bases. Those metallic oxides which are bases are called *basic oxides*.

All metallic hydroxides are bases, because with acids they give salts and water only. The following is a list of some metallic hydroxides with their corresponding oxides. They are all bases.

Metallic hydroxide	*Corresponding oxide*
Caustic soda (sodium hydroxide), NaOH	Sodium oxide, Na_2O
Caustic potash (potassium hydroxide), KOH	Potassium oxide, K_2O
Slaked lime (calcium hydroxide), $Ca(OH)_2$	Calcium oxide, CaO
Magnesium hydroxide, $Mg(OH)_2$	Magnesium oxide, MgO
Ammonium hydroxide, NH_4OH	No corresponding oxide
Cupric hydroxide, $Cu(OH)_2$	Cupric oxide, CuO

ALKALIS

We have noticed that some substances turn red litmus blue, and we have called some of them alkalis, but it would be wrong to call all of them alkalis. The term alkali is used very loosely in popular language, and even chemists are not always strict in its use. The popular idea is that an alkali is something that is the 'opposite' of an acid. The idea is good so far as it goes, but we must try to get a clearer idea than that.

Most metallic oxides will not dissolve in water. Some will. Sodium oxide and potassium oxide readily combine with water to form the corresponding hydroxides which are very soluble in water.

$$Na_2O + H_2O = 2NaOH.$$

$$K_2O + H_2O = 2KOH.$$

Calcium oxide (quicklime) dissolves in water and forms the slightly soluble calcium hydroxide.

$$CaO + H_2O = Ca(OH)_2.$$
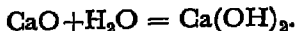
Magnesium oxide and lead monoxide dissolve in water to a slight extent, forming the hydroxides of the metals, which are soluble in water. All metallic hydroxides are not soluble in water, but when they are, their solutions turn red litmus blue. They all have the property of neutralizing an acid, that is, they take away the acid properties from the acid, at the same time losing their own alkaline properties. They do this by forming salts and water, and nothing else.

An *alkali* may be defined as a soluble hydroxide. There are four very important and common alkalis: caustic soda, caustic potash, slaked lime, and ammonium hydroxide. We have already studied slaked lime. We shall leave the study of ammonium hydroxide until we have studied ammonia. Caustic soda and caustic potash are so much alike that it will be enough if we study one of them. The two cannot be distinguished by their appearance, and in most of their chemical actions they are alike.

CAUSTIC SODA

Expt. 171. Examine a piece of caustic soda. The pure substance is usually made in the form of sticks and is kept in sealed bottles. Remove a portion of a stick and place it in a clean dish where you can examine it. It is almost white and apparently non-crystalline. Touch it with the finger. It feels wet and soapy. Wash your hands after handling it otherwise the skin may come off where it has been in contact with the caustic soda, hence the name caustic (burning) soda. After about ten minutes note the amount of solution in the dish. Caustic soda is *deliquescent*, i.e. it takes up so much water from the air that it can dissolve in it.

Drop a little solid caustic soda in a test-tube containing a little water. Note how quickly it dissolves. Feel the bottom of the test-tube as the caustic soda dissolves. It is warm. When caustic soda dissolves in water, heat is evolved. Test the solution with litmus paper. It turns red litmus paper blue. Put another piece of paper in a strong solution of caustic soda. The paper is destroyed, turning into a pulp.

Expt. 172. To neutralize hydrochloric acid with caustic soda. To a solution of caustic soda in an evaporating basin add dilute hydrochloric acid, and stir. Remove a drop of the liquid in the basin on a glass rod and put it on a piece of litmus paper. By this means you can tell if the solution is still alkaline or acid. Keep adding the acid a little at a time, stirring all the time, until the liquid is just acid. If you add too much by mistake, add a little caustic soda solution and start all over again. With patience you will be able to obtain a solution which is almost neutral, that is, neither acid nor alkaline. If it were exactly neutral it would not turn the colour of either red or blue litmus, but it is very difficult to reach this stage, so we arrange that it shall be slightly acid rather than slightly alkaline. What have you in your solution now? It must be a salt and water only.

$$NaOH + HCl = NaCl + H_2O.$$

Evaporate the solution to dryness and you will obtain common salt.

Do you see why we had the solution slightly acid rather than alkaline? It was because the slight excess of hydrochloric acid is driven off on evaporation, but an excess of caustic soda is not driven off.

This is a common method of making salts. Most sodium salts can be prepared by neutralizing caustic soda solution with the appropriate acid. We shall use this method again with other acids and other alkalis.

QUESTIONS

1. What relationship exists between acids, bases, and salts? What is an acid salt? What classes of substances are bases? Illustrate your answers by reference to actual substances.

2. What do you understand by the terms: (a) water of crystallization, (b) efflorescence, (c) deliquescence? Give examples of each.

3. What is an alkali? Give the names and formulae of four important alkalis. Name a salt which has an alkaline reaction.

4. Describe in detail how you would make a dry sample of copper sulphate crystals starting from cupric oxide (CuO).

5. Calculate the percentage weight of water of crystallization in washing soda.

6. What weight of sodium chloride is obtained when 5 gm. of caustic soda is neutralized by hydrochloric acid?

XIV

HYDROCHLORIC ACID AND THE HALOGENS

HYDROCHLORIC ACID

Expt. 173. To a little common salt (sodium chloride, NaCl) in a test-tube add a few drops of concentrated sulphuric acid. Feel the bottom of the test-tube. What has caused it to become hotter? Note the gas fuming at the mouth of the test-tube. Blow across the mouth of the test-tube. The gas fumes more. Cautiously smell the gas. It has a pungent odour.

Hold a glass rod which has been dipped in ammonium hydroxide solution in the fumes. They change to dense white fumes which look bluish in some lights.

Put a lighted splint in the gas. The gas does not burn and the light is put out.

Note the substance left in the test-tube.

The colourless gas, fuming in the air, that is obtained is hydrochloric acid gas. It is commonly called spirits of salt because it is a 'spirit', i.e. a gas, which comes from common salt. The formula of its molecule is HCl, and its chemical name is hydrogen chloride.

$$NaCl + H_2SO_4 = NaHSO_4 + HCl.$$
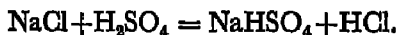
The substance left in the test-tube is sodium hydrogen sulphate, or sodium bisulphate, an acid salt of sulphuric acid. If the experiment had been performed under conditions which allowed the substances to be strongly heated, we should have had more hydrogen chloride evolved and the normal sodium sulphate left.

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$
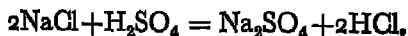
Expt. 174. Repeat the above experiment using, instead of sodium chloride, other chlorides such as potassium chloride, KCl, calcium chloride, CaCl_2, and ammonium chloride, NH_4Cl. In every case hydrogen chloride is evolved, and a sulphate, either normal or acid, is obtained.

PREPARATION AND PROPERTIES OF HYDROGEN CHLORIDE

Expt. 175. Use the same apparatus as in Fig. 115. Pour concentrated sulphuric acid down the thistle funnel on to some common salt in the flask. Collect the gas by downward delivery. When the action slows down, instead of adding more acid, warm the flask gently. In this way the rate of the evolution of the gas can be controlled. Collect several jars of the gas. Test the gas with a lighted splint. Add a few drops of ammonium hydroxide solution to a jar of the gas. Add water to another jar of the gas, shake, and add blue litmus. Hydrogen chloride is very soluble in water and this property may be demonstrated as follows:

Expt. 176. Fill a dry, round-bottomed flask with the gas by downward delivery. Have ready a cork which exactly fits the flask, and through which passes a long glass tube drawn out to a nozzle at the end which goes into the flask, and fitted with an india-rubber tube and clip at the other end (Fig. 118). Before fitting the tube to the flask, fill the tube with water by allowing water to flow through it from the clip end, and by applying the clip as the water is still flowing. Fit the cork to the flask full of hydrogen chloride. Place the clip end of the tube in water coloured with blue litmus solution. Release the clip while it is under water.

Water turns red as the acid dissolves in it,

Water coloured blue with litmus
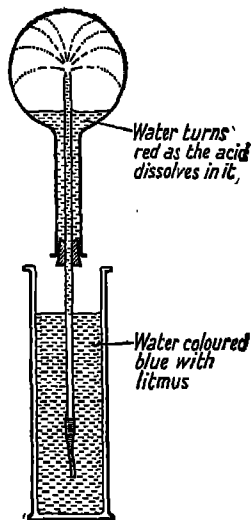
FIG. 118. To show the great solubility of hydrochloric acid gas in water.

The drop of water at the nozzle dissolves some of the hydrogen chloride, thus decreasing the pressure in the flask. A little more water therefore enters the flask and dissolves more of the gas. Suddenly the water enters the flask through the nozzle at such a rate that a fountain is formed, the litmus turning from blue to red on entering.

This fountain experiment is only possible with a gas which is extremely soluble in water, because it depends on a sudden reduction of pressure in the flask being caused by most of the gas dissolving in the water.

A solution of hydrogen chloride in water is called hydrochloric acid or spirits of salt. The gas is the real compound hydrogen chloride, and the substance called hydrochloric acid is a solution. This is unlike sulphuric acid which is a liquid itself and not a solution. It is necessary to take precautions when dissolving a very soluble gas in water. An arrangement for so doing is shown in Fig. 119. If the water dissolves the gas so quickly that the water rises up the funnel, there is no danger of the water reaching the flask where the gas is being made, because the level of the water outside the funnel will drop, thus letting in the air and restoring the pressure.
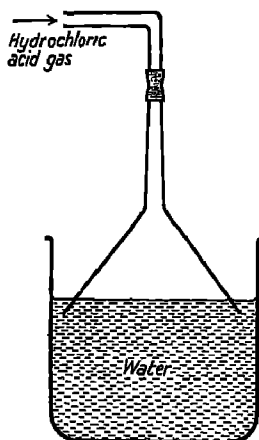
Fig. 119. Dissolving hydro-chloric acid gas in water.

Expt. 177. Make a solution of the gas and compare its action with that of dilute hydrochloric acid from a reagent bottle, on litmus, sodium carbonate, zinc, and silver nitrate solution. The solutions have the same properties.

When a few drops of hydrochloric acid are added to silver nitrate solution a white precipitate of silver chloride is obtained, which darkens on exposure to light.

$$HCl + AgNO_3 = AgCl + HNO_3.$$
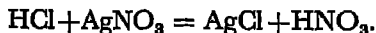
This is a useful test for the presence of hydrochloric acid or a soluble chloride.

Silver chloride is used with silver bromide for making photographic sensitive films.

When ammonia gas, NH_3, comes in contact with hydrogen chloride, dense white fumes are formed. These fumes consist of large numbers of tiny particles of a solid called ammonium chloride, NH_4Cl. It is formed by the combination of ammonia and hydrogen chloride:

$$HCl + NH_3 = NH_4Cl.$$
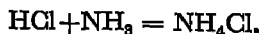
We shall have more to say about this compound when we study the chlorides.

The chief properties of hydrogen chloride are as follows: gaseous, colourless, fumes in the air, heavier than the air, extremely soluble in water, puts out a lighted splint, does not burn itself, turns blue litmus solution red, and gives dense white fumes with ammonia. We have seen also that hydrochloric acid will react with some metals to form chlorides, hydrogen being evolved in the process.

Hydrochloric acid is used for preparing chlorine. It is present in a very dilute form in the gastric juices of many animals, where it plays an important part in the process of digestion (Ch. XLIV).

SOME SALTS OF HYDROCHLORIC ACID

Salts of hydrochloric acid are called *chlorides*. If the hydrogen of hydrochloric acid is in some way replaced by a metal, then a chloride is formed. We shall now consider some of the commoner salts of the acid.

1. Sodium chloride, NaCl

This salt is so common that it is called *common salt*, or more frequently just *salt*. It occurs in enormous quantities in the sea, but the evaporation of sea water does not yield pure sodium chloride because other salts such as potassium chloride and magnesium chloride occur in sea-water. It also occurs in enormous quantities underground. In some parts of the world it is mined as *rock salt*. In Cheshire, where the main salt beds of Britain occur, the salt is no longer mined because water has found its way to the salt and mining is impossible. It is easy, however, to pump to the surface the water which has been in contact with the salt and has dissolved it. This is done, and the *brine* (salt solution) evaporated when the salt is required.

This removal of the salt from under the ground has caused the surface of the ground to subside in many places, notably in Northwich. In this town many of the buildings are out of the perpendicular. Some houses have had to be abandoned because of this.

Common salt is used extensively in chemical industries, and many of these in England are carried on not far from the salt district in Cheshire, e.g. Runcorn and Widnes.

Common salt such as is used in cooking or at the table is not pure sodium chloride. It usually contains a little calcium sulphate and the chlorides of calcium and potassium. *Table salt* is required finely grained and must not be deliquescent. The fine grain is obtained by evaporating the brine very quickly near the boiling-point. Removal of the chlorides of magnesium and calcium prevents the salt from being deliquescent. Salt used for preserving fish is required in large crystals. This form of salt is obtained by slow evaporation.

Sodium chloride occurs in the blood and is required in the food of all animals. Carnivorous animals (those which eat other animals) obtain the salt in the blood of the animals they eat. Herbivorous animals obtain their salt from the plants they eat. Some also go to salt springs and to places in which salt rock appears near the surface, often travelling great distances to do this.

Human beings obtain their salt from the food they eat, both animal and vegetable; but it seems that the supply is not sufficient in most cases. For this reason people who live in a country containing little salt and far from the sea have to import it. Bags of salt are carried long distances in Africa, and some tribes even use salt as a means of exchange, like money.

In countries near the sea, the salt is obtained by the evaporation of sea-water. This is easy in tropical and sub-tropical countries, but difficult in Arctic regions. On the shores of the Mediterranean the sea-water is allowed to flow into very shallow ponds which are then cut off from communication with the sea. The hot sun evaporates the water and the salt crystallizes out. It is removed with shovels perforated to allow the water to drain away. The wet salt is left on the ground by the side of the ponds to be dried by the sun.

In the Arctic the salt water of the sea is concentrated by freezing. When salt water of a certain strength partially freezes, the ice first formed is almost pure water. Consequently the sea-water left is a more concentrated solution of salt. This

concentrated solution is further concentrated by the heat from wood fires. (It must not be thought that a salt solution of *any* concentration will give pure ice on freezing. Above a concentration of 23·6 per cent. of salt, salt separates out on freezing. At a concentration of 23·6 per cent. both salt and ice freeze together as one mass.)

The question is often asked: Why is the sea salt? We have given one possible answer to this question on p. 13. The question, however, might well be answered by saying: Why should it not be salt? When sodium chloride was formed from sodium and chlorine as the earth cooled, it would presumably be more or less evenly distributed, and it would naturally occur in that part now occupied by the sea. It would also occur on the land, but as it is very soluble in cold water, much of it would be gradually washed into the sea.

How did the salt deposits on land arise? Most of them were almost certainly deposited during the evaporation of a portion of the sea cut off from the main sea by some land movement. Further land movements, upheavals and subsidences, account for the present underground position of the beds. We must not be too sure, however, because these things happened a long time ago, and it is sometimes very difficult to decide what has happened in a particular case.

2. Potassium chloride, KCl

This salt is very much like sodium chloride. If you had a pure specimen of both salts, you could not distinguish between them by their appearance. If a little potassium chloride is held on a platinum wire, or on a gas 'mantle' rod, in the almost colourless flame of a bunsen burner, a violet colouration is imparted to the flame by the potassium salt. If this is done with sodium chloride, a yellow colour is imparted to the flame. Platinum, or a mantle rod, is used because it does not affect the flame. If there is any sodium chloride present as an impurity in the potassium chloride, this test will not work because the yellow flame, due to the sodium, masks the violet flame, due to the potassium.

Potassium chloride is not so soluble as sodium chloride in

cold water, but in hot water it is much more soluble than sodium chloride. The latter is very little more soluble in hot water than in cold (p. 282).

3. Calcium chloride, CaCl₂

This salt is obtained as a by-product during many manufacturing processes. It is a white, deliquescent solid, often used for drying gases and in desiccators. The gas to be dried is passed through a U-tube containing pieces of calcium chloride. Of course, the method is useless for those gases which react with calcium chloride, such as ammonia.

4. Ammonium chloride, NH₄Cl

This substance is commonly known as SAL-AMMONIAC. It is formed when ammonia gas comes in contact with hydrochloric acid. It is a white solid, soluble in water. It is used as a fertilizer, and in *Leclanché* and *dry cells*. There is no compound which has the composition which would justify giving to it the formula NH₄, and yet this group of elements often occurs in compounds just as if the group were a metal. For example, we look upon the compound ammonium chloride as if it were a salt of hydrochloric acid in which the hydrogen of the acid has been replaced, not by a metal, but by the group of elements NH₄ (called the *ammonium group*) acting as a metal. We shall meet other compounds later in which this group occurs, and in every case it will be acting chemically as if it were a metal.

When ammonium chloride is heated it changes into a vapour without first turning into a liquid. On cooling it changes directly into solid ammonium chloride, again omitting the liquid stage. This process is called *sublimation*.

Expt. 178. Heat a little ammonium chloride in a test-tube. After a short time note the solid ammonium chloride deposited on the cooler parts of the test-tube. Between the solid at the bottom of the test-tube and the solid at the top, note the clear space which is occupied by the vapour which turns into ammonium chloride on cooling.

CHLORINE

Expt. 179. The action of concentrated hydrochloric acid on lead peroxide. Cover a little brown oxide of lead (lead peroxide) in a test-tube with concentrated hydrochloric acid. Warm gently. Note the pale green gas given off. Smell it cautiously. It has a choking odour. Test with a lighted splint. The gas puts out the light and does not burn itself. Dip a piece of filter paper in ink and push it into the mouth of the tube. The ink disappears. Do the same with wet litmus paper. It turns white.
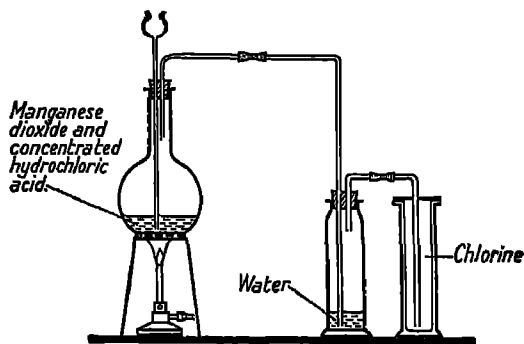
FIG. 120. Preparation of chlorine.

The gas is chlorine. The lead peroxide reacts with the hydro-chloric acid to give lead chloride, water, and chlorine.

$$PbO_2 + 4HCl = PbCl_2 + 2H_2O + Cl_2.$$

Here we have an oxide of a metal reacting with an acid to give not only a salt and water, but also something else. The metallic oxide is therefore not a base. It is a *peroxide*. Peroxides always contain more oxygen than the corresponding basic oxide. In this case the extra oxygen has combined with some hydrogen from the hydrochloric acid, thus setting free chlorine.

The experiment should be repeated using manganese dioxide instead of lead peroxide. Manganese dioxide is a peroxide and chlorine is obtained in the same way. This is the way chlorine is usually prepared in the laboratory.

PREPARATION AND PROPERTIES OF CHLORINE

Expt. 180. To prepare several jars of chlorine. Prepare the gas in a fume cupboard as chlorine is poisonous. Fit up the apparatus as shown in Fig. 120. Cover the granulated manganese dioxide in the flask with concentrated hydrochloric acid. Warm gently. The gas passes through the wash bottle of water and is collected by downward delivery, chlorine gas being very much heavier than air. The wash bottle is to free the chlorine of hydrochloric acid gas which comes off with the chlorine. The hydrochloric acid gas is very soluble in water and remains dissolved in the water of the wash bottle. Chlorine is less soluble in water, and although some dissolves, much more passes through the water. As the experiment proceeds, note the gradual spreading of the green colour from the flask, through the wash bottle, and into the jar. When a jar is full, change it and collect several jars. Chlorine can also be prepared by adding concentrated hydrochloric acid to solid potassium permanganate. The acid must enter the flask drop by drop as the action is vigorous. No heat is required.

From our experiments we already know some properties of the gas. It is pale green, has a choking smell, is heavier than air, is soluble in water, puts out a lighted splint, and does not burn.

Test the solubility of the gas in caustic soda solution by inverting a test-tube of the gas over it. Chlorine passed into caustic soda solution forms sodium hypochlorite. This is used as a disinfectant, and is the main constituent of many commercial preparations used for this purpose.

Expt. 181. Into a jar of the gas pour some litmus solution. The colour disappears. Put other moistened coloured substances, e.g. red cloth, flower petals, ordinary writing ink marks on paper, printed matter, &c., in the gas. Some of them are bleached, but others are not, e.g. printer's ink.

Chlorine will not bleach when the substances containing the dye are dry. The chlorine dissolves in the water and some of it reacts with the water to form hydrochloric acid. The oxygen thus released oxidizes the dye of the material to a colourless substance. This is not the full explanation, but it is accurate so far as it goes.

Expt. 182. Put a lighted wax taper into a jar of the gas. It burns

with a smoky flame. Now dip a glass rod into ammonia solution and hold it over the mouth of the jar. Observe the dense white fumes of ammonium chloride showing the presence of hydrogen chloride.

The wax of the taper is a mixture of hydrocarbons. The chlorine combines with the hydrogen to form hydrogen chloride while the carbon is set free and appears as smoke.

Expt. 183. Burn hydrogen from a jet in a jar of chlorine (Fig. 121). Follow the same method and *take the same precautions* as in Experiment 129. Test for the presence of hydrogen chloride as before.

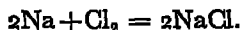The hydrogen burns to give hydrogen chloride.

$$H_2 + Cl_2 = 2HCl.$$
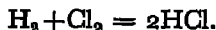
A mixture of hydrogen and chlorine in equal proportions by volume will explode when exposed to bright sunlight, forming hydrochloric acid gas. In diffused sunlight the combination takes place quietly and gradually.

Fig. 121. The burning of hydrogen in chlorine.

Chlorine affects the mucous membranes of the nose, throat, and lungs. One good breath of chlorine is dangerous. If you breathe a small quantity of chlorine and become affected, a quick but cautious smell of ammonia from the bottle gives relief and prevents further trouble, but ammonia is no remedy in serious cases.

Chlorine was the first gas to be used by the Germans in the Great War (1914–18). The box-respirator is effective against it because chlorine is a dense gas (about two and a half times as dense as air) and all dense gases are absorbed by the carbon in the box-respirator (p. 177).

Expt. 184. Warm gently a *small* piece of sodium in a deflagrating spoon until it melts and place it in a jar of chlorine. It burns with the production of sodium chloride.
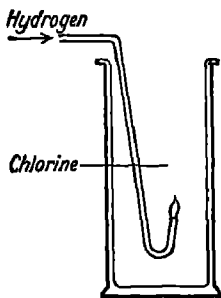
$$2Na + Cl_2 = 2NaCl.$$

This is proof that common salt is a compound of sodium and chlorine.

COMMERCIAL PREPARATION OF CHLORINE

Chlorine is no longer prepared commercially by the action of hydrochloric acid on manganese dioxide. This method was superseded by the Deacon process in which the hydrochloric acid is oxidized by the oxygen of the air with the aid of a catalyst. Most chlorine to-day is manufactured by decomposing sodium chloride in solution by electrolysis (Ch. XL). Chlorine is produced, and the sodium set free reacts with the water to form caustic soda and hydrogen. Thus three substances of great commercial importance are manufactured in one operation.

BLEACHING POWDER ('CHLORIDE OF LIME')

Much of the chlorine produced commercially is used in the manufacture of bleaching powder. When chlorine is brought into contact with dry slaked lime, the chlorine is absorbed and bleaching powder, or 'chloride of lime', is produced.

$$Ca(OH)_2 + Cl_2 = CaOCl_2 + H_2O.$$
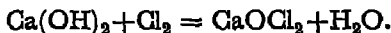
Expt. 185. Examine some bleaching powder. Note the smell. It is a little different from that of chlorine in a jar. Add dilute acid to a little bleaching powder in a test-tube. Chlorine is evolved. Dip some coloured fabric in an aqueous solution of bleaching powder until it is thoroughly soaked. Transfer the solution to a weak solution of acid. The fabric is bleached.

By this method the chlorine is liberated in the fibres of the fabric and the bleaching occurs there. If the cloth is now washed in water and then with a solution of sodium thiosulphate ('hypo') the last traces of chlorine are removed. The cloth can now be washed with water again and dried.

Cotton and linen goods can be bleached by chlorine but not silk or wool, as they are injured by the action of the chlorine.

Bleaching powder is also used as a decontaminator to destroy the poison gas known as mustard gas.

USES OF CHLORINE

Chlorine, usually in the form of bleaching powder, is used extensively in bleaching. Besides being used for bleaching fabrics it also bleaches wood-pulp from which paper and

artificial silk are made. It is also used for sterilizing bathing and drinking water, as chlorine, even in small quantities, kills harmful bacteria.

The water in many public plunge baths is not renewed. Instead it is regularly run out of the bath, filtered through gauze, aerated, and then 'chlorinated' on its return to the bath, usually by the addition of a few drops of liquid chlorine. Bleaching powder may also be used for this purpose.

Some public water authorities now sterilize drinking water by the use of chlorine, and means have been found to eliminate the taste given to the water by the chlorine. There is, however, still a prejudice against the sterilization of drinking water by this method. In times of war, when it is impossible to be too particular about the source of drinking water, 'chlorination' of the water by the use of bleaching powder is a great safeguard against water-carried diseases.

Some public health authorities give free, to any householder who asks for it, a quantity of bleaching powder under the name of 'chloride of lime'. It is intended to be used in drains and around sinks as a disinfectant. When exposed to moist air the carbonic acid formed by the water and the atmospheric carbon dioxide acts on the bleaching powder like other dilute acids and releases chlorine. This, as we have seen, is a good disinfectant. To many people a faint smell of chlorine is not unpleasant. They have learned to think of it as a 'clean smell'.

Chlorine is used in the manufacture of many dyes and drugs and the poison gases known as phosgene and mustard gas. It is also used in the preparation of carbon tetrachloride, a fire extinguisher, and the anaesthetic chloroform.

THE HALOGENS

There are three other elements which have properties similar to those of chlorine: fluorine, bromine, and iodine. These four elements form the class called *halogens*, a Greek word meaning 'salt-producing'. Chlorine is the commonest of the halogen elements and we have dealt with it and its chief compounds in some detail. We shall not deal with fluorine, but something must be said about bromine and iodine.

IODINE is a deep violet, almost black, solid with a metallic lustre. It is, however, a non-metal like chlorine and bromine. When heated slowly it melts and then changes into a beautiful violet vapour. When heated quickly it changes directly into the vapour without passing through the liquid state, and the vapour changes into the solid on cooling. In other words, it *sublimes* (cf. ammonium chloride), forming small, shiny, nearly black crystals.

Expt. 186. To show that iodine sublimes. Place a little iodine in a hard glass beaker, and on the latter place an evaporating dish containing cold water. Heat the iodine quickly. Note the violet vapour which changes to solid iodine on the underside of the evaporating dish. Do not heat longer than necessary or the beaker will break.

Iodine is insoluble in water, but it dissolves in an aqueous solution of potassium iodide. It will also dissolve in alcohol. The liquid commonly called iodine is either a solution of iodine in potassium iodide solution, or the 'tincture of iodine', which is an alcoholic solution of iodine containing some potassium iodide. The liquid is often used for putting on cuts to prevent them from turning 'septic'. The solution in potassium iodide is used as a testing reagent for starch (p. 126).

Iodine is present in a compound known as thyroxin in the thyroid glands of animals. A deficiency of iodine in the diet causes serious diseases, for then this gland cannot function properly. 'Iodized salt' is ordinary table salt to which has been added a very small amount of an iodine compound. Many people regularly use this salt in order to prevent a deficiency of iodine in the diet.

Iodine occurs in sea-water in the form of sodium iodide and to a larger extent in sea plants. Seaweed is collected by the inhabitants of the western shores of Scotland, Ireland, and France, and burned. The resulting ash (called *kelp* in Scotland) contains a small proportion of iodine chiefly as sodium and potassium iodides. The kelp is treated with concentrated sulphuric acid and manganese dioxide, resulting in the liberation of iodine vapour which condenses to the solid form.

Much of the iodine used in commerce is obtained from the mother liquor during the extraction of sodium nitrate (p. 252).

BROMINE, like chlorine and iodine, is not found free in nature but in combination with metals. It is present in the salts crystallized out from sea-water, and also occurs in salt deposits and in certain sea plants. It is a heavy, reddish-brown liquid. The vapour has an irritating smell and attacks the mucous membranes of the nose and throat. It is poisonous and causes sores on the skin. Its solution in water, called *bromine water*, is often used in the laboratory instead of the element itself.

Bromine is used in the manufacture of certain dyes. Potassium bromide is used extensively as a sedative in medicine. Silver bromide is very sensitive to light and is used for coating photographic films and printing papers.

Chlorine, bromine, and iodine closely resemble one another in their properties. Chlorine displaces bromine and iodine from solutions of bromides and iodides respectively, and bromine displaces iodine but not chlorine. Iodine will not displace either of the other two. The following table illustrates the intermediate nature of bromine.

THE PROPERTIES OF CHLORINE, BROMINE, AND IODINE COMPARED

Property	Chlorine	Bromine	Iodine
State	Gas	Liquid	Solid
Colour	Green	Reddish-brown	Almost black
Melting-point	$-102°$ C.	$-7·3°$ C.	$114°$ C.
Boiling-point	$-34°$ C.	$60°$ C.	$184°$ C.
Reaction with hydrogen	Combines with hydrogen at ordinary temperatures	Combines with hydrogen when heated	Combines with hydrogen with difficulty when heated
The silver salt	Silver chloride. White precipitate, soluble in ammonia	Silver bromide. Yellow precipitate slightly soluble in ammonia	Silver iodide. Yellow precipitate insoluble in ammonia

All these precipitates are insoluble in dilute nitric acid. They all darken in daylight due to reduction to silver.

QUESTIONS

1. Starting from common salt, how would you make some dilute hydrochloric acid? Write an equation for the reaction. Where does this acid occur in the human body?

2. Give the chemical names and formulae of (a) spirits of salt, (b) sal-ammoniac. Describe what happens when sal-ammoniac is heated. Give one common use of sal-ammoniac.

3. You are given four pure salts and told that one is common salt, another potassium chloride, a third sal-ammoniac, and the fourth calcium chloride, but they are unlabelled. State exactly how you would distinguish between them.

4. What is common salt? State how it is obtained and give its chief properties and uses. What kind of salt is desirable as a table salt? What is iodized salt and what is its use?

5. How is chlorine prepared in the laboratory and how is it manufactured on a large scale? What are its uses and in what forms is it obtained commercially?

6. Compare the chief properties of chlorine, bromine, and iodine, and give the sources and common uses of the last two elements and their compounds. If you ask for iodine at a pharmacist's what will you be given?

7. What volume of chlorine, measured at N.T.P., is obtained when sufficient manganese dioxide is heated with 5 gm. of hydrochloric acid?

XV

NITRIC ACID AND AMMONIA

NITRIC ACID

Expt. 187. The action of concentrated sulphuric acid on some nitrates. Cover a little potassium nitrate (also called *nitre* and *saltpetre*) in a test-tube with a little concentrated sulphuric acid. Watch what happens in the cold. Note that the bottom of the test-tube becomes warmer. The potassium nitrate partly dissolves in the acid. Warm gently and note very carefully what happens. The salt dissolves in the acid, and a vapour is given off which condenses into an oily, colourless liquid on the cooler part of the test-tube. Do not heat strongly or this vapour will be driven out of the test-tube and damage your fingers. Test the vapour with a lighted splint. Hold a piece of damp litmus paper at the mouth of the test-tube. Hold the test-tube against a light background and note the brown fumes in it. Allow the test-tube to cool. What is the appearance of the substance left at the bottom of the test-tube? Repeat the experiment with sodium nitrate. Apparently the same vapour is evolved.

The vapour given off is nitric acid. At the temperature of the reaction the acid is a vapour, but at the temperature of the room it is a liquid. The pure vapour is colourless and so is the pure liquid. The brown colour is due to a gas called nitrogen peroxide. This is caused by the decomposition of some of the nitric acid when heated. When concentrated sulphuric acid acts on a nitrate, nitric acid is formed:

$$KNO_3 + H_2SO_4 = KHSO_4 + HNO_3.$$
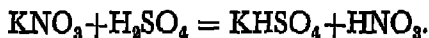
The potassium in the potassium nitrate is replaced by the hydrogen of the acid. The substance left in the test-tube is potassium hydrogen sulphate, the acid sulphate of potassium. The action of sulphuric acid on nitrates may be compared with the action of that acid on chlorides. In both cases the acid of the salt is produced. This is the usual action of concentrated sulphuric acid on a salt.

THE PREPARATION AND PROPERTIES OF NITRIC ACID

Expt. 188. Fit up the apparatus illustrated in Fig. 122. Use potassium nitrate or sodium nitrate in the retort, and add concentrated sulphuric acid until the salt is well covered. Heat gently at first and never so strongly that much of the vapour is driven out of the retort. As the vapour travels down the neck of the retort it is condensed. The liquid is collected in the flask kept cool by cold
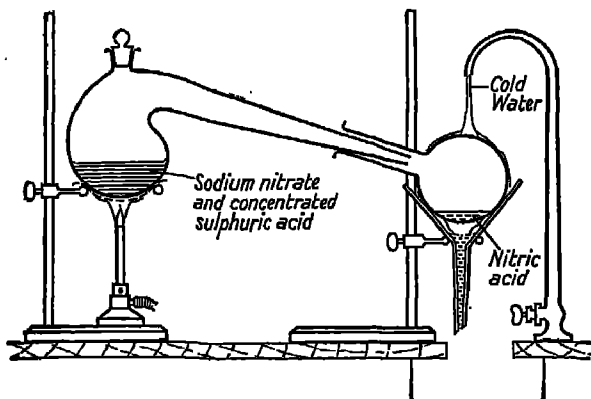

FIG. 122. Preparation of nitric acid.

water. Note the change in colour as the temperature increases. The liquid obtained is not colourless but brownish-yellow. This, and the colour of the vapour, is due to some of the acid decomposing and the resulting nitrogen peroxide dissolving in the acid or mixing with the vapour. When no more nitric acid is evolved, remove the bunsen burner and carry the retort to a fume cupboard. Remove the stopper of the retort and pour the liquid contents into an evaporating dish. Allow to cool and then examine it.

Pure nitric acid is an oily, colourless liquid which is easily vaporized. As usually seen in the laboratory it is slightly yellow due to dissolved nitrogen peroxide, NO_2. The acid is easily decomposed into nitrogen peroxide, oxygen, and water:

$$4HNO_3 = 4NO_2 + 2H_2O + O_2.$$

The decomposition is brought about by heat or the presence

of sunlight. Hence, if nitric acid is to be kept colourless it should be kept in the dark.
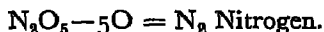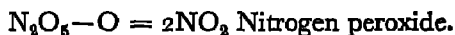
The Action of Nitric Acid on some Metals

In our study of sulphuric and hydrochloric acids, we saw that the usual action of the acid on the metal was to produce a salt and hydrogen, the metal taking the place of the hydrogen of the acid and so forming a salt. In the case of nitric acid the salt is formed as usual, but hydrogen is rarely evolved. This is because nitric acid, besides being an acid, is also a strong oxidizing agent. It readily parts with some of its oxygen to oxidize anything that will combine with oxygen. If nitric acid acted as an acid only, hydrogen would be evolved; but it acts also as an oxidizing agent and the hydrogen is oxidized to water before it can appear. In oxidizing the hydrogen to water, the acid must lose some oxygen. Let us see what will be formed when nitric acid loses oxygen.

Nitric acid may be looked upon as nitrogen pentoxide, N_2O_5, combined with water. Nitric anhydride is nitrogen pentoxide.

$$N_2O_5 + H_2O = 2HNO_3.$$

When nitric acid loses oxygen for oxidation purposes it is the nitric anhydride that is reduced and not the water. This may be effected as follows:

$$N_2O_5 - O = 2NO_2 \text{ Nitrogen peroxide.}$$
$$N_2O_5 - 2O = N_2O_3 \text{ Nitrogen trioxide.}$$
$$N_2O_5 - 3O = 2NO \text{ Nitric oxide.}$$
$$N_2O_5 - 4O = N_2O \text{ Nitrous oxide.}$$
$$N_2O_5 - 5O = N_2 \text{ Nitrogen.}$$

All these products are possible, but the commonest is nitrogen peroxide.

It is possible in one case to obtain hydrogen when nitric acid acts on a metal. When dilute nitric acid acts on magnesium in the cold, some hydrogen is usually evolved.

Expt. 189. To a little cold, very dilute nitric acid in a test-tube add a small piece of magnesium ribbon. Test the gas for hydrogen

by holding a lighted splint in the mouth of the test-tube. What happens to the magnesium? The equation for the action is:

$$Mg + 2HNO_3 = Mg(NO_3)_2 + H_2.$$
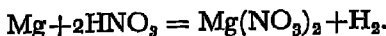
This action is similar to the action of dilute sulphuric and hydrochloric acids on magnesium.

Expt. 190. The action of concentrated nitric acid on magnesium. Put a little concentrated nitric acid in a test-tube, which must not be held in the hand, but placed in a test-tube stand. Drop into the acid a piece of magnesium ribbon. There is a violent reaction, the magnesium dissolves, and there is an evolution of the reddish-brown gas nitrogen peroxide.

$$Mg + 4HNO_3 = Mg(NO_3)_2 + 2NO_2 + 2H_2O.$$
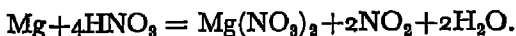
Here the nitrate is again formed but in this case the hydrogen is oxidized to water and the nitric acid is reduced.

Nitric acid will act on most metals. The concentrated acid usually acts as above, but when the acid is dilute, nitric oxide, NO, or nitrous oxide, N_2O, may be evolved instead of nitrogen peroxide.

Expt. 191. Try the action of concentrated nitric acid on copper and zinc. Follow the instructions given in Experiment 190.

Nitric acid is used for dissolving metals. In etching, a design is drawn on a waxed sheet of copper which is then placed in a bath of nitric acid. This 'eats' into the copper where the wax has been removed. The result is that a design is etched into the copper and can be reproduced on sheets of paper.

Most metals are attacked by nitric acid and it was formerly called *aqua fortis*—the strong water—for this reason. It will not dissolve gold. If a substance looks like gold but will dissolve in nitric acid it may safely be said that it is not gold. This is called the 'acid test'. Gold can be dissolved in a mixture of concentrated hydrochloric acid and concentrated nitric acid in the proportion of 3 : 1 by volume. This mixture is called *aqua regia*—the kingly water—because it dissolves the 'king' of metals, gold. Platinum is another metal which is not attacked by nitric acid but which dissolves in aqua regia.

Dilute nitric acid stains the skin yellow, and the strong acid causes painful sores. It attacks organic matter.

Expt. 192. Try the action of concentrated nitric acid on wood, paper, and india-rubber. Drop a little of each in separate test-tubes containing the acid. Most of the wood disappears, only a few fibres being left; the paper dissolves; and the rubber disappears very quickly. In each case nitrogen peroxide is given off, so it would appear that something is being oxidized.

The explosives gun cotton, nitroglycerine (the active part of dynamite), picric acid, and T.N.T. (tri-nitrotoluene) are all made by the action of nitric acid on organic substances.

SOME IMPORTANT SALTS OF NITRIC ACID

1. Potassium nitrate, KNO_3 (saltpetre or nitre)

We used this nitrate in the preparation of nitric acid. It can be prepared in the laboratory by neutralizing caustic potash with dilute nitric acid, concentrating the liquid by evaporation, and then allowing the crystals to separate out on cooling. It is a white, crystalline solid, soluble in water. It does not become damp on exposure to the air. When heated strongly it decomposes into oxygen and potassium nitrite:

$$2KNO_3 = 2KNO_2 + O_2.$$
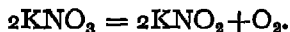
If a small piece of carbon is dropped into a test-tube containing boiling molten nitre, it burns brilliantly, the oxygen required for combustion being that liberated from the nitre.

'Touch paper' is made by soaking strips of absorbent paper in a solution of potassium nitrate and then drying the paper. The paper readily smoulders if ignited.

Potassium nitrate is a constituent of gunpowder. This is a mixture of potassium nitrate, sulphur, and carbon. When ignited the products of the action are chiefly solid potassium sulphide, nitrogen, and carbon dioxide. The last two are gases, and occupy a volume greatly in excess of that occupied by the original solids. The volume is further increased by the enormous rise of temperature in the reaction (p. 437). Hence if the gunpowder is ignited in a closed space, the gases exert a great pressure which is used for propelling or shattering objects. (You should not attempt to make gunpowder or experiment with it. It is *dangerous*, besides being *illegal*.)

2. Sodium nitrate, $NaNO_3$ (Chile saltpetre)

This substance is very much like potassium nitrate. It is a white, crystalline solid, much more soluble in water than the potassium salt, and deliquescent. For this reason it is unsuitable as a component of gunpowder. If it were used, the gunpowder would become wet on exposure to the air and thus become useless. On heating sodium nitrate decomposes like potassium nitrate.

$$2NaNO_3 = 2NaNO_2 + O_2.$$

Sodium nitrate occurs in immense deposits in South America, chiefly in Chile, hence its name. All nitrates are soluble in water, consequently the natural deposits are found in places of little rainfall. The rock bearing the nitrate is blasted and the sodium nitrate extracted with water. It is exported in great quantities, and is used as a fertilizer and in the manufacture of nitric acid.

3. Ammonium nitrate, NH_4NO_3

This substance is used in the preparation of nitrous oxide, N_2O. On heating the nitrate decomposes into nitrous oxide and water:

$$NH_4NO_3 = N_2O + 2H_2O.$$

Nitrous oxide is sometimes called 'laughing gas', because when inhaled it often causes hysterical laughter. It is the gas used by dentists and surgeons for producing a period of unconsciousness while an operation is being performed. Formerly the gas was used by itself, but it is now administered mixed with oxygen, and can thus be used with safety for longer periods of unconsciousness. It is a very useful anaesthetic for this purpose, and hysteria is not normally produced when the gas is administered under these conditions.

The salt ammonium nitrate is not an explosive, but there is sometimes an explosion if, in the preparation of nitrous oxide, the heating is carried on until there is very little substance left in the flask. It is safer to stop the action while there is still some ammonium nitrate not decomposed.

The salt is used in the manufacture of explosives. Mixed with another substance it is the explosive in the 'Mills' bomb.

It is not used as a fertilizer because it is more expensive than the other nitrogen fertilizers.

AMMONIA, NH₃

This compound, like nitric acid, contains nitrogen. Its pungent odour can often be detected when organic matter
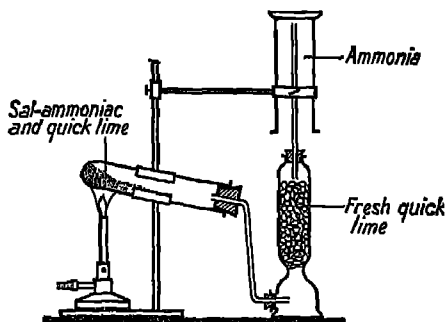

FIG. 123. Preparation of ammonia.

decays or burns, e.g. near stables and in blacksmiths' forges when the hot iron shoe is applied to the hoof of a horse. It used to be prepared by the dry distillation of the horns and hoofs of animals, and was known as 'spirits of hartshorn'. It is produced when any ammonium compound is heated with an alkali. In the laboratory it is conveniently prepared by the action of slaked lime on sal-ammoniac:

$$2NH_4Cl + Ca(OH)_2 = CaCl_2 + 2H_2O + 2NH_3.$$
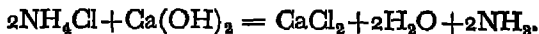
Expt. 193. To prepare and collect ammonia. Mix some powdered sal-ammoniac with some dry slaked lime in a mortar. Note the smell of ammonia as the substances are mixed. The gas is produced even in the cold. Hold over the mixture a piece of damp red litmus paper. It turns blue.

Put some of the mixture in a dry boiling tube sloping downwards towards the mouth, and fit up the apparatus as shown in Fig. 123. Heat gently. Water is formed during the action but when it condenses it runs to the cooler part of the tube and does not crack it. The gas is

dried by passing it up a drying tower containing quicklime. The usual drying agents, sulphuric acid and calcium chloride, cannot be used because ammonia reacts with them. Collect the gas by upward delivery, as ammonia is less dense than air. By holding a glass rod which has been dipped in hydrochloric acid at the mouth of the jar you will be able to judge when the jar is full of the gas. Fill a flask with the gas and perform the 'fountain' experiment as you did with hydrogen chloride. This time let the water running into the flask be coloured with red litmus (p. 233).

PROPERTIES OF AMMONIA

Ammonia is a colourless gas, less dense than air, and extremely soluble in water. The saturated solution has a relative density of 0·88. When a trace of ammonia is inhaled it is not unpleasant, but it takes very little to bring tears to the eyes. When breathed in any quantity it is dangerous.

**Expt. 194.** Put a lighted taper into a jar of the gas. The light is extinguished and the gas does not burn.

If dry ammonia gas is passed through a heated glass tube drawn out to a jet, the gas will burn. The heat causes the gas to decompose into its elements, nitrogen and hydrogen, and it is the latter which burns:

$$2NH_3 = N_2 + 3H_2.$$
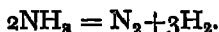
Even without heating the tube the gas can be made to burn if the lighted taper be held to the jet all the time. This means that ammonia is very easily decomposed by heat.

Expt. 195. Pass a slow current of oxygen through a 0·88 solution of ammonia in a beaker. Heat a spiral of thin platinum wire and hold it in the mixture of gases above the liquid in the beaker. The wire continues to glow.

Ammonia will burn in oxygen.

$$4NH_3 + 3O_2 = 2N_2 + 6H_2O.$$
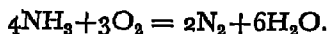
The gas turns damp red litmus paper blue, but the dry gas does not change the colour of dry litmus paper. When the gas dissolves in water some of it combines with the water to form ammonium hydroxide, NH_4OH.

$$NH_3 + H_2O = NH_4OH.$$
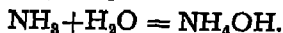
We have seen that ammonium hydroxide is a base, and with

acids gives rise to ammonium salts. Ammonia gas, however, is not a base, because with an acid it gives a salt but no water. It is the hydroxide that is the base. Ammonia gives dense white fumes with hydrochloric acid. We have seen that these fumes consist of ammonium chloride.

Ammonia dissolved in water is used for cleaning purposes as it will remove most kinds of household grease. Much of the ammonia manufactured is used in the manufacture of ammonium sulphate and ammonium chloride. Ammonium carbonate is used in smelling salts and 'sal volatile'.

Ammonia gas is very easily liquefied by pressure. When liquid ammonia is allowed to evaporate, the heat necessary is taken from the surroundings which are thus cooled (p. 454). Ammonia is used extensively for making ice, keeping refrigerators cold, and freezing the water of indoor skating rinks.

NITROGEN FROM THE AIR

All plants need nitrogen. This they obtain from the soil as soluble nitrogen compounds. With very few exceptions plants are unable to use the free nitrogen of the air. Every crop harvested removes nitrogen from the soil, and if this nitrogen is not replaced in some way the soil will become less fertile. This is a very important question and we must discuss it in some detail later (Ch. XLVII). Farmers replace the nitrogen in the soil by dressings of farm-yard manure, but the supply of this is limited, and it has to be supplemented by artificial fertilizers containing nitrogen, such as ammonium sulphate and sodium nitrate. The supply of sodium nitrate (Chile saltpetre) is enormous, but so is the world demand. It was seen that unless some other source of nitrogen was found, the supply of Chile saltpetre would eventually be exhausted, and the world would be faced with the fact that the soil was becoming less able to support the increasing population.

In looking round for future nitrogen supplies it was natural that the possibility of using the nitrogen of the air should be considered, a problem which has come to be known as the *fixation of nitrogen.*

The first method employed on a large scale was to convert

the nitrogen and oxygen of the air to nitric oxide. This is possible at the high temperature of the electric arc (about 3,000° C.).
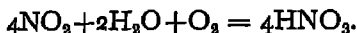
$$N_2 + O_2 = 2NO.$$
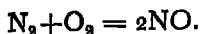
Only a small fraction of the oxygen and nitrogen, however, combines to form nitric oxide.

The nitric oxide so formed combines with more oxygen of the air to form nitrogen peroxide:

$$2NO + O_2 = 2NO_2.$$
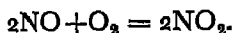
The nitrogen peroxide was then made to combine with water and more oxygen of the air to form nitric acid:

$$4NO_2 + 2H_2O + O_2 = 4HNO_3.$$

From the nitric acid obtained any nitrate can be made.

Germany, cut off during the First World War (1914–18) and again during the Second World War from supplies of nitrate from Chile, was able to manufacture nitrate for high explosives by this process.

This method is not now in common use, and is in fact only used in countries where cheap electric power is available from waterfalls. It has been superseded by a process in which the nitrogen of the air is made to combine directly with hydrogen to form ammonia. It is known as the *Haber process*. Nitrogen from the air and hydrogen are heated under pressure in the presence of a metallic catalyst. Some of the hydrogen and nitrogen combine. The gases are then cooled and the ammonia condenses to a liquid, the uncombined gases going through the process again. Eventually most of the hydrogen and nitrogen is converted into ammonia, from which any nitrogen compound required can be prepared.

Heat is required only to start the action. This process, unlike the one it has superseded, is not dependent on a cheap supply of heat.

Free nitrogen in the air is of no commercial value; combined nitrogen is valuable. We are no longer dependent on the deposits of nitrates in South America, and the fear of a world nitrogen shortage has disappeared.

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QUESTIONS

1. Describe briefly how you would prepare nitric acid in the laboratory. How does the action of this acid on magnesium differ from that of hydrochloric acid on the same metal? Explain, giving equations.

2. Write short notes on each of the following: (*a*) the 'acid test', (*b*) aqua fortis, (*c*) aqua regia, (*d*) saltpetre, (*e*) Chile saltpetre.

3. Explain, giving an equation, why concentrated nitric acid kept in the laboratory is slightly yellow. How can it be kept colourless?

4. Describe how nitric acid is used in etching. State any other uses of nitric acid.

5. What is the action of heat on: (*a*) potassium nitrate, (*b*) ammonium nitrate? Give equations. What use is made of the gas evolved in (*b*)?

6. If you were given any ammonium salt how would you get ammonia gas from it? How is this gas usually prepared in the laboratory? What tests would you make to find out if a colourless gas was ammonia? State the uses of ammonia.

7. What do you understand by the term 'fixation of nitrogen'? Why is the process necessary and how is it performed?

8. What weight of nitric acid is required to dissolve completely 1 gm. of magnesium?

9. What volume of ammonia (measured at N.T.P.) is evolved when 5 gm. of ammonium sulphate is heated with sufficient slaked lime? What weight of slaked lime is sufficient for this purpose?

METALS AND ALLOYS

METALS AND ALLOYS IN THE SERVICE OF MAN

EVERYONE has some general idea of what a metal is like and no doubt could name a dozen without much difficulty. Our present civilization depends on the use of metals. Think, for a moment, what the world would be like without metals. Our present civilization could not have arisen if we had been confined to the use of stone and wood as men were in the Stone Age. Think of transport without metals. What kind of houses could be built without metallic tools and fittings? What kind of a world would it be without iron and steel? We need not stress the importance of metals to man.

Man, in his search for metals suitable for his purposes, soon found that metals in the pure state were seldom the best, and he soon learnt to add elements to them in order to improve their qualities for particular purposes. Indeed, it is probable that the metals first extracted by primitive man were not pure and for that reason were more suited to his purposes.

If a metal is blended with another metal the metallic substance obtained is called an *alloy*. For example, when copper and zinc are blended the alloy brass is obtained. The term alloy is now also used when the second element added is not a metal. Thus steel is now considered to be an alloy of iron and carbon, although the latter element is a non-metal.

The properties of the alloys are often quite different from those of their constituent elements and pure metals are now seldom used in the construction of buildings and machines as alloys with more suitable properties can be employed. Man can now choose from the metals and the alloys a metallic substance suitable for almost any purpose.

Some metals and alloys have been used by man since very early times. Primitive man made weapons, tools, and utensils of copper, brass, and bronze. Iron implements, probably

5,000 years old, have been found, and copper has been used since prehistoric times. Copper occurs uncombined in isolated places and perhaps man thus obtained his first supplies, but he learnt also to extract it from its ore. The same is probably true of iron. The production of iron by primitive man was a great event, for iron was a much better metal than copper for making weapons and tools, and those tribes in possession of iron had great advantages over those who had weapons and tools made only of copper.

The Romans, at the time of the invasion of Britain, employed iron, bronze, copper, brass, lead, tin, silver, and gold. The last two were called the Noble Metals, partly because they were little affected by the atmosphere and water, and also because they could be melted time after time with little loss. The Romans at the time of the invasion used bronze swords and some of their coins were made of brass.

As we study some of the commoner metals and their alloys we shall try to understand how man first came to use them, how he obtained them from their ores, and what improvements have been made in modern times on the methods of primitive man.

SOME REACTIONS OF METALS

We have already learnt something about metals in our previous work. We have experimented with potassium, sodium, calcium, magnesium, iron, zinc, copper, and lead, and studied their reactions with oxygen, with water, and with acids. We have seen that a salt is formed when the hydrogen of an acid is replaced by a metal and that some oxides of metals react with water to form alkalis. All these reactions should be revised before proceeding.

THE CHARACTERISTIC PROPERTIES OF METALS

Elements are classified into metals and non-metals. When all the elements are taken into consideration the classification is not easy and a satisfactory definition of a metal is not known. Some elements, e.g. gold and copper, are undoubtedly metals,

others, e.g. carbon and sulphur, are undoubtedly non-metals; but some, e.g. arsenic and antimony, partake of the character of both metals and non-metals. There is a gradual transition through the elements from those having metallic properties to those having the properties of non-metals.

Let us consider the chief characteristics of metals and non-metals. We shall study only the usual properties of metals and non-metals and it must be remembered that there are important exceptions.

All metals, except mercury, are solids at the ordinary temperature, and non-metals, with the exception of bromine, are either solids or gases. Of course, this is only a question of temperature. Mercury at the north pole would be a solid, but it is usual to find that metals have to be raised to a high temperature before they melt and to a much higher temperature before the liquids so formed boil. Non-metals, on the other hand, usually have low melting- and boiling-points.

Metals are usually dense, that is, for their volume they are very heavy. A cubic foot of lead, for example, is much heavier than a cubic foot of carbon or sulphur. The gaseous non-metals are, of course, very 'light'.

The appearance of an element may tell us whether it is a metal or not. Metals have a *lustre*, that is, a shining glossy surface, while non-metals are 'earthy' in appearance. We cannot, however, depend on this quality as graphite and iodine, both non-metals, have a metallic lustre and might easily be mistaken for metals if we judged only by appearance.

A very important property of metals is that they are usually strong and tough. They withstand much pulling without breaking whereas non-metals usually have very little *tensile strength* (i.e. they cannot withstand much pulling without breaking). Some metals have more tensile strength than others; a steel wire, for example, supports a greater weight than does a lead wire of the same dimensions.

Metals can also be hammered and beaten into various shapes and into thin sheets without cracking. They are *malleable*. Gold is the most malleable of all metals. You have probably seen 'gold leaf'. A property allied to this is *ductility*.

Metals are ductile, that is, they can be stretched and otherwise deformed without cracking. Non-metals are not malleable or ductile: they are *brittle*.

Metals are good conductors of heat and electricity while non-metals are not usually so. We shall learn much more about this in Chapter XXXVII.

Dinner gongs and bells are made of metallic substances because they give out a rich or powerful sound when struck. They are *sonorous*.

So far we have dealt only with the physical properties of metals, and these are very important when we are thinking of metals for building and engineering; but the chemical properties are also important, not only to chemists, but to the engineer, for he has to consider the effects on his materials of exposure to air, water, and other liquids. The chemical properties of a metal also affect the extraction of the metal from its ore.

Some metals oxidize easily and some with difficulty, but every metal forms at least one basic oxide (e.g. cupric oxide, CuO) whereas the oxides formed by non-metals are not basic. They are either acidic, like carbon dioxide, CO_2, or neutral, like carbon monoxide, CO. Again, metals do not easily combine with hydrogen to form hydrides, and when they do the compounds are usually unstable and unimportant. Non-metals, on the other hand, easily combine with hydrogen forming very important and usually stable hydrides, e.g. ammonia, NH_3, sulphuretted hydrogen, H_2S, and methane, CH_4.

The chlorides of metals are true salts, e.g. sodium chloride, $NaCl$, and they are not usually decomposed by water. The chlorides of the non-metals, on the other hand, are not salts. They are usually decomposed by water into hydrochloric acid, HCl, and an acid of the non-metal. Thus phosphorus trichloride is decomposed by water into hydrochloric acid and phosphorous acid.

We shall now study some common metals and their alloys, and see how the above characteristics affect their extraction and the uses to which man puts them.

IRON

Iron was discovered and used very early in the history of man but not so early as copper. It is possible that free (i.e. uncombined) iron was the first source. It is certain that isolated tribes in modern times have used the free iron in meteorites which occasionally reach the earth from outer space. All that was necessary was to hack off pieces of iron as they were needed. This source, however, is very small and the great bulk of the iron used by primitive people must have been obtained from compounds of iron. Iron is acted on by the atmosphere, and it is a fairly active element so it is not found free in nature except in a few isolated places.

Probably the first extraction was accidental. Some iron ore, an oxide, was perhaps heated in a wood charcoal fire, and when the fire died down the solid iron was found. The chemical reaction is simple. It is merely the reduction of the oxide, carbon being the reducing agent.

$$Fe_2O_3 + 3C = 2Fe + 3CO.$$

The iron obtained in this primitive way was what we now know as *wrought iron.* It was not melted during the process and the primitive men were able to bend and hammer it into the shapes they required. No doubt they soon found that they could do the shaping much better by hammering and bending while the iron was hot. Very soon the extraction was not accidental. They deliberately reduced the iron ore with wood charcoal and they probably increased the heat by using bellows.

We have seen that iron does not occur free in nature except in small quantities, but compounds of iron are very widely distributed. All soils contain a little and many rocks are coloured by compounds of iron (Ch. XLVII). Green plants cannot grow without iron, and it enters into the composition of the haemoglobin of the blood. In some parts of the world, e.g. the Urals of the U.S.S.R., there are very large deposits of iron ore. An ore is a mineral consisting of a compound of the metal often mixed with earthy matter. If the metal is not easily extracted the source is not considered an ore.

The chief iron ores are haematite, Fe_2O_3, and spathic ore, $FeCO_3$. When the latter is heated in the air it decomposes into the oxide. To obtain iron from the ore the oxide has to be reduced to the metal. We have seen that this was done in primitive times. In the fourteenth century the operation was carried out in furnaces, the reducing agent being wood charcoal. A blast of air was blown into the furnace. Under such conditions the iron melts and when solidified is known as *cast iron*. It contains about 5 per cent. of carbon and a little sulphur and phosphorus. It is brittle and cannot be shaped by hammering.

Much wood was necessary for the production of the necessary charcoal so that those ores which were found adjacent to forests were exploited. Much deforestation occurred before a new method of extraction was discovered using coke instead of wood charcoal. Then those ores found near the coalfields were exploited. At present there is little production of ore in England, nearly all of it being imported.

The modern extraction of iron is an improvement on the early method but it is still a reduction process. It is carried out in a blast-furnace which is a tall tower sometimes 100 feet tall made with fire-resisting brick. The ore is first roasted in the air. This changes any carbonate to the oxide and also drives off any water which may be combined with the oxide. The ore is then mixed with coke and limestone and the mixture fed into the furnace at the top. A blast of air, pre-heated to about 700° C., enters the furnace at the bottom. The carbon of the coke combines with the oxygen of the air in the blast to form carbon monoxide:

$$2C + O_2 = 2CO.$$
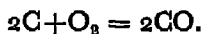
The carbon monoxide acts as a reducing agent, reducing the iron oxide to iron.

$$Fe_2O_3 + 3CO = 3CO_2 + 2Fe.$$
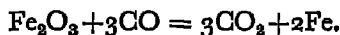
The molten iron sinks to the bottom of the furnace.

The earthy matter in the ore would react with the molten

iron in the blast-furnace if limestone were not present. The limestone is changed to calcium oxide in the furnace:

$$CaCO_3 = CaO + CO_2.$$

The calcium oxide reacts with the earthy matter and forms a molten slag which floats on the surface of the iron at the bottom of the furnace.

The molten iron is run off periodically. The slag floats on the top and is raked off before the iron runs into moulds made of sand. The solidified iron at this stage is called *pig iron* because of the likeness of the channels in the sand to a sow with her litter.

The hot gases which ascend the furnace consist largely of carbon monoxide. They are used to heat the air before it is passed into the blast-furnace.

Cast iron contains about 5 per cent. of carbon with a little sulphur, phosphorus, and silicon (Ch. XLVII). It is brittle both when cold and hot and consequently cannot be hammered, drawn, or welded. It is, however, very fluid when molten and can be cast into very intricate moulds. It is used as castings for cylinders, stoves, radiators (not for car radiators), desk fittings, lathe beds, 'chairs' for railway lines, water pipes, and for many other things not likely to be subject to shock.

Wrought iron is nearly pure iron containing only about 0·06 per cent. of carbon and a little slag. It is made by removing almost all the carbon from cast iron. The latter is melted in a reverberatory furnace (i.e. one where the heat is reflected from the top on to the surface of the metal). The floor of the furnace is made of haematite, the oxygen in which reacts with the carbon in the molten cast iron to form carbon monoxide,

$$Fe_2O_3 + 3C = 2Fe + 3CO,$$

while the silicon, phosphorus, and sulphur are oxidized. The sulphur dioxide escapes and the oxides of silicon and phosphorus form a slag. As masses of the metal begin to solidify they are removed and hammered and squeezed to remove the slag.

Wrought iron has a higher melting-point than cast iron and is easily softened by heat when it can be shaped and welded by

heating. It is used for horseshoes, chains, fire-bars, gates, railings, and nails. Formerly it had many other uses but it has been largely replaced by steel.

Steel always contains carbon but the amount varies according to the quality of steel desired. There is a general purpose steel which contains as much as 0·2 per cent. of carbon but others contain from 0·09 to 0·1 per cent. As steel contains more carbon than wrought iron and less than cast iron it can be made (i) by adding carbon to wrought iron, (ii) by removing carbon and impurities from cast iron and then adding the exact amount of carbon required, and (iii) by mixing cast iron with wrought iron. All these processes are carried out at the present day but the details are beyond the scope of this book.

By including the correct amount of carbon (and sometimes small quantities of other elements) and by appropriate treatment, a steel can be made hard, or soft, or ductile to almost any degree so that it is possible to manufacture a steel suitable for any particular purpose. Varieties of steel are used in ships' plates, building construction, railway lines, and most things formerly made from wrought iron.

Invar is another alloy of iron. It contains 36 per cent. of nickel. It is very useful because it expands and contracts very little with change of temperature (p. 416). It is used for making accurate measuring rods and for parts of accurate clocks.

COPPER

The use of copper is of great antiquity. Next to iron it is the most useful metal. Although it was probably first found free it was later extracted from its ores, perhaps accidentally in the first place. Some ores of copper are oxides and carbonates and they have only to be reduced. Bronze, an alloy of copper and tin, was probably first discovered in some part of the world where copper and tin ores were found together, and brass, an alloy of copper and zinc, might have been similarly discovered by smelting copper and zinc ores together. As bronze was found to be hard and useful no doubt the ores were hence-

forward deliberately smelted together. Both copper and bronze were in use in Britain before Roman times.

The ores used at present are copper pyrites, $CuFeS_2$, and the oxides, carbonates, and sulphates. The carbonates are first converted into the oxide by roasting in the air and the oxide is then reduced with carbon in a reverberatory furnace. The copper is extracted from the sulphates electrolytically (Ch. XL), but the method of extraction from copper pyrites is complicated. The copper is purified by electrolysis (Ch. XL).

Copper is reddish-brown and has a very attractive appearance. Near its melting-point it is brittle and can be powdered. When heated and then cooled slowly it is brittle, but with rapid cooling it becomes soft, malleable, and ductile. It resists corrosion. The green substance found on copper after exposure to the air is known as verdigris. It is the basic carbonate of copper, i.e. a compound of the carbonate and the hydroxide. A coat of verdigris resists further action of the atmosphere.

Copper is (after silver) the best conductor of electricity (Ch. XXXVII). To have its highest conductivity the copper must be pure, and more than half of the pure copper produced is used in the electrical industry. Copper is a very good conductor of heat and is used in locomotive fire-boxes, domestic water boilers, and petrol pipes. It is used in the radiators of cars because, in addition to its being a good conductor of heat, it is also easily brazed and soldered.

Bronze is an alloy of copper and tin. The more tin there is in the alloy the harder it is. Bronze was probably the earliest alloy made and primitive man varied the amount of tin to produce different bronzes for his purposes. Our 'copper' coinage is a familiar bronze containing besides copper 0·5 per cent. of tin and 2·5 per cent. of zinc. Gun metal, not now used for making guns, is a bronze of composition 88 per cent. copper, 10 per cent. tin, and 2 per cent. zinc. Aluminium bronze contains copper with 10 per cent. of aluminium. It is three times as strong as copper. Manganese bronze containing about 3 per cent. manganese is used for the propellor blades of steamships. It is tougher than ordinary bronze and is not much corroded by sea-water. German silver contains from 56 to 60 per cent.

copper, 20 per cent. zinc, and 20 to 25 per cent. nickel. It is used for imitating silver and in the construction of resistance coils.

Brass is another valuable alloy of copper. It usually contains about 70 per cent. copper and 30 per cent. zinc. The brass cartridges used in rifles are of this composition. When the brass contains over 36 per cent. of zinc it is hard and strong, but below this percentage brass is ductile. Brass buttons are made of a brass which is ductile enough to be pressed.

ALUMINIUM

We have seen that iron and copper with some of their alloys have been in use since prehistoric times. Aluminium, on the other hand, has come into general use only during the last fifty years. This is not because aluminium is a rare metal. On the contrary, it is very widely distributed. It occurs in granite and therefore in clays, because when granite weathers clays are formed. It is, however, very difficult to extract aluminium from clay and even to-day no economical method of extraction from clay has been discovered. The chief aluminium ore is bauxite, $Al_2O_3.2H_2O$, but this oxide is not easily reduced as are those of iron and copper. Much energy is required and the electrolytic method is employed. It is for this reason that aluminium is a 'new' metal. We deal with the extraction of aluminium by electrolysis in Chapter XL.

Aluminium is a silvery metal with a beautiful lustre. It is about one-third as light as steel and is a very good conductor of electricity. It is also a good conductor of heat. It resists corrosion because a very fine film of oxide protects it from further action.

Aluminium is now in common use. The pure metal is used for cooking utensils and in the chemical industry. It is the wrapping foil known as 'silver paper'. The overhead electric conductors stretching from pylon to pylon across the country are made of aluminium with a central core of galvanized steel wire to give the conductor the necessary tensile strength (Ch. XLII).

Aluminium forms alloys of high tensile strength. *Duralumin*

is a trade name given to a series of aluminium alloys which contain about 3·5 per cent. of copper and 0·5 per cent. of magnesium. They are light and strong. For a given weight, duralumin is more than twice as strong as steel. It is therefore used extensively in aircraft framework where great strength in relation to weight is required. As duralumin is rather corrodible it is usually covered with pure aluminium which resists corrosion. It is then called 'Alclad'.

ZINC

Zinc was not isolated until many hundreds of years after the discovery of brass. Zinc is volatile (i.e. easily vaporized) and probably escaped into the air when early attempts were made to isolate it. The chief ores are zinc blende, ZnS, and calamine, $ZnCO_3$. When these are roasted both give the oxide:

$$ZnCO_3 = ZnO + CO_2,$$
$$2ZnS + 3O_2 = 2ZnO + 2SO_2.$$
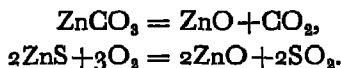
The extraction is then a simple process of reduction as with iron and copper. The oxide is mixed with powdered coke and heated in retorts. The zinc is given off as a vapour which is condensed to a liquid by condensers fixed to the retorts.

$$ZnO + C = Zn + CO.$$

Much zinc is now extracted by the electrolytic process (Ch. XL).

The uses of zinc depend on the fact that the metal is non-corrosive. Iron or steel is often covered with zinc (by dipping into molten zinc) and this combines the strength of steel with the non-corrosive property of zinc. A 'galvanized iron' bucket is made of mild steel and coated with zinc. Galvanized iron is often used for kitchen fittings, for roofings, or for any other purpose involving protection from the weather. Much zinc is used in dry batteries (Ch. XXXVII) and in the manufacture of brass. We dealt with this as an alloy of copper (p. 267).

Zinc oxide is a white powder obtained when zinc is heated in the air. It is yellow when hot and white when cold. It is usually made by heating the carbonate. It is used in medicine (zinc

oxide plaster) and as a basis for a white paint (zinc white). It is not so good as white lead paint but has the advantage over lead in that it does not darken with age (p. 221).

TIN

Tin occurs naturally as the oxide SnO_2. The ore is known as cassiterite. It is found chiefly in Malay, Bolivia, and Mexico. Tin mines have been worked in Cornwall from ancient times. It is extracted by reducing the oxide to the metal. The ore is first crushed in running water. The heavy oxide sinks to the bottom and much of the earthy matter is washed away. The ore is then roasted in the air and this removes more earthy matter especially sulphur which escapes as the oxide. The purified oxide is then reduced with anthracite in a reverberatory furnace:

$$SnO_2 + 2C = Sn + 2CO.$$
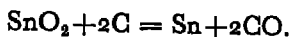
Tin is non-corrosive and its chief use is in making tinned plate. Ordinary 'tin' utensils such as 'tin' cans are made of tinned plate which is mild steel dipped in molten tin. The utensils are strong like steel but do not rust until the tin has worn off. Until aluminium came into general use tin was used as a 'silver paper' wrapping.

LEAD

We need only look around to see what a useful metal lead is. It has also useful alloys and compounds. Lead has been known for a very long time, probably over 2,000 years. Its chief ore is galena, PbS. This is easily converted into the oxide by heating in air and then the oxide is reduced. No special reducing agent is required as some of the galena can be used for this purpose.

$$2PbO + PbS = 3Pb + SO_2.$$
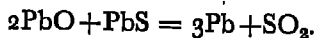
Lead has a low melting-point, is very heavy, weak, and soft. It can easily be moulded into shape when cold or cast into moulds when liquid. It is used for making pipes, lining gutters, and lining roofs. When exposed to the atmosphere it quickly

acquires a very thin coating of oxide which acts as a protective coat preventing further corrosion. If you scratch lead which has been exposed to the air you will see the bright metal exposed. This quickly tarnishes. Lead is attacked by soft water (p. 272) in the presence of air, but lead pipes can be safely used for ordinary household water as this is usually slightly hard and a coating is formed on the inside of the pipe which prevents further action. As lead compounds are poisonous lead pipes could not be used for household water if the lead was attacked by it.

Solder is a useful alloy of lead. It is made by mixing lead and tin together when the metals are molten. Various qualities of solder can be made by varying the amount of tin added. Plumber's solder is usually about half lead and half tin. It has the useful property of gradually turning solid from liquid. At 220° C. the liquid begins to solidify but it is not hard until it has cooled to about 180° C. Because of this it is very useful for joining pieces of lead together, especially for connecting lead pipes. It gives the plumber time to mould the solder. Watch a plumber 'wiping a joint'. He melts the solder with a blow-pipe and wipes with a cloth the pasty solder round the ends of the pipes to be joined. He could not do this with pure lead or pure tin because they solidify immediately without going pasty.

Type metal is an alloy containing about 60 per cent. lead, 30 per cent. antimony, and 10 per cent. tin. It is hard, thin when molten, and expands slightly on solidifying thus taking a sharp impression of the mould. The solid metal contracts on cooling.

Some compounds of lead are in common use. Litharge, PbO, is a yellow powder used in making lead glazes for pottery and also as a 'dryer' in paint. Red lead, Pb_3O_4, is used in making good-quality red paint. White lead is the basic carbonate, $2PbCO_3.Pb(OH)_2$. It makes an excellent paint but it is affected by sulphur compounds in the atmosphere and turns dark owing to the formation of lead sulphide, PbS, which is black (p. 221).

QUESTIONS

1 Write notes for an essay on 'Man's use of Metals and Alloys'.

2. Distinguish, as far as you can, between metals and non-metals. What (*a*) chemical properties, (*b*) physical properties of metals are important to building engineers and machine designers?

3. How was iron obtained in primitive times? Describe briefly how it is now obtained from its ores giving equations for any chemical reactions you mention.

4. Distinguish between wrought iron, cast iron, and steel. Show how their uses are connected with their physical and chemical properties.

5. What is an alloy? Name *one* alloy of *each* of the following metals and give their properties and uses: iron, copper, nickel, zinc, lead, aluminium.

6. Iron is a very 'old' metal and aluminium a comparatively 'new' metal. Account for this. State the properties and uses of duralumin.

7. What are the uses of tin? Why is it suitable for these uses? What other metals may be used for the same purpose?

8. What is solder? Compare its properties with those of the metals from which it is made. Why is it suitable for repairing burst lead pipes? Describe this operation.

9. What is (*a*) white lead, (*b*) zinc white, (*c*) red lead? What are their properties and uses?

XVII

THE HARDNESS OF WATER

IT is now possible to undertake a further study of water as a solvent, because we know something of the chemistry of the substances which are commonly found dissolved in natural waters.

You know from your experience that it is more difficult to

COMPARISON OF THE HARDNESS OF VARIOUS SAMPLES OF WATER

Sample of water	No. of c.c. of soap solution used for		Diff. between (i) and (ii)
	Unboiled sample (i)	Boiled sample (ii)	
Distilled			
Tap			
Sol. of cal. bicarb.		·	
Sol. of mag. sulph.			

wash in some waters than in others. We find that with some water it is difficult to get a lather with soap, while with other water a lather comes almost immediately. We say that water is *hard* when it will not easily give a lather with soap; we call it *soft* when it will easily give a lather. We must try to discover what causes some waters to be hard, and how to remove the hardness.

It is possible to compare samples of water with reference to hardness. We do this, in fact, when we wash our hands in different waters, but we must use a more exact method if we wish our comparisons to be anything more than rough estimates. Instead of using soap it is found more convenient to use soap solution, which enables us to measure the amount used more accurately.

To prepare a soap solution. Cut some white windsor soap into thin shavings (or use soap flakes). Put them into a bottle and add cold distilled water. Shake the bottle well until the soap has dissolved. Such a solution is suitable for the experiments we have to do. There is no need to use a standard solution, or one of any particular strength. It is important, however, that the same solution, without further additions of water, be used for all the experiments.

Expt. 196. To compare the hardness of some samples of water. Use the following samples: (*a*) distilled water; (*b*) tap-water; (*c*) a solution of calcium bicarbonate (made by passing carbon dioxide through clear lime-water until the cloudiness that first forms has disappeared); (*d*) a solution of a little magnesium sulphate in distilled water.

Transfer 25 c.c. of the sample to a titration flask by means of a pipette (Fig. 124). Run in from a burette 1 c.c. of soap solution. Shake. If a lather is formed, leave it for a few minutes to see if it remains. A lather which remains for a few minutes is called a 'permanent lather'. If the lather is not permanent, add another c.c. of soap

FIG. 124. Measuring vessels used in titration.

solution and shake. Repeat until a permanent lather is obtained. How many c.c. of soap solution were required to give a permanent lather? Record your results in a table as shown above.

Now take another 25 c.c. of the same sample, but before adding soap solution boil the sample for a few minutes and then let it cool. Now continue as before and record your results in the same table.

TEMPORARY AND PERMANENT HARDNESS

Consider first the unboiled samples. Distilled water requires very little soap solution before a permanent lather is obtained. Tap-water is harder; the solution of calcium bicarbonate much harder; and a solution containing even a moderate amount of magnesium sulphate is very hard.

What is the effect of boiling the solutions? It makes no difference to the distilled water. It was soft before boiling and it remains soft after boiling. It makes no difference to the magnesium sulphate solution. This was hard before boiling and remains hard after boiling. With tap-water the boiling does make a difference. The water is softer after boiling, but it has not lost all its hardness. Again, with the calcium bicarbonate solution boiling removes the hardness.

There are two kinds of hardness of water, namely: *temporary hardness*, which is removed by boiling, and *permanent hardness*, which is not removed by boiling.

CAUSES OF THE HARDNESS OF WATER

Dissolved calcium and magnesium salts cause water to be hard. Suspended matter, even suspended salts of calcium and magnesium, does not cause hardness. The compounds commonly causing hardness are calcium bicarbonate, $Ca(HCO_3)_2$, magnesium bicarbonate, $Mg(HCO_3)_2$, calcium sulphate, $CaSO_4$; magnesium sulphate, $MgSO_4$, calcium chloride, $CaCl_2$, and magnesium chloride, $MgCl_2$.

When a solution of calcium bicarbonate is boiled, insoluble calcium carbonate is formed:

$$Ca(HCO_3)_2 = CaCO_3 + H_2O + CO_2.[1]$$

Similarly insoluble magnesium carbonate is formed when a solution of magnesium bicarbonate is boiled. Therefore any

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[1] This is an example of a *reversible reaction*, for if carbon dioxide is passed into a suspension of chalk in water calcium bicarbonate is formed. The equation representing this reaction (a *reversible equation*) often has the sign ⇌ instead of the usual sign. Other examples of reversible reactions are the action of iron on steam, and mercury on oxygen.

hardness due to the presence of the bicarbonates of calcium and magnesium is removed on boiling, because the calcium and magnesium salts are removed from solution. They are still in suspension, but suspended matter does not cause hardness.

The sulphates and chlorides of calcium and magnesium are not altered on boiling, and the hardness due to these salts is not removed.

Expt. 197. Add a little soap solution to some hard water and shake.

At first there is an appearance of a lather, but as the solution settles it is seen that there is a white, insoluble scum on the surface of the water. This scum represents wasted soap. Let us see what soap is and how it acts on water.

SOAP

Expt. 198. Warm small quantities of various fats and oils, e.g. lard, suet, olive oil, linseed oil, with caustic soda solution and shake well. Notice whether the oil or fat is dissolved and whether a lather is formed on shaking.

Dissolve a very small piece of caustic soda in a few c.c. of methylated spirits and add about half the quantity of a fat or oil, and warm on a water bath. Pour the contents of the tube into a strong solution of common salt. Separate the solid curd and rinse it with water. See if it will produce a lather with water.

Soap is made from natural animal and vegetable fats. The fats are boiled with caustic soda for ordinary hard soap, and with caustic potash for soft soap. Glycerine is also formed. Excess of alkali is used so that all the fat will be changed. In good soaps all the alkali left over is carefully removed as it has an irritating effect on the skin and a destructive action on cloth; but in some cheap soaps the alkali is not completely removed. This free alkali may easily be detected by testing the soap solution (in distilled water) with red litmus.

Ordinary soap is the sodium salt of a 'fatty acid'. The process of heating the fat with caustic soda is known as *saponification*. The soap is 'salted out' by pouring the product into a strong solution of salt, since soap is insoluble in this strong solution. It is filtered and pressed, the glycerine being thus separated from it.

As soft soap cannot be 'salted out' it contains glycerine and other substances.

Expt. 199. Shake up a little olive oil with water, and compare with the result obtained when olive oil is shaken up with soap solution. Allow them both to stand.

The milky liquid obtained by shaking oil and water is called an *emulsion*, which means 'like milk'. The oil breaks up into tiny globules which are suspended in water. The milkiness quickly disappears and the two liquids separate out again. If some substance is added which prevents the globules from joining together again, a true emulsion is formed. Milk is itself an emulsion.

One reason why soap and water together make a good cleansing agent is that the soap makes a thin film round the particles of dirt to be removed. They are thus much more easily washed away. Another and probably more important reason is that soap forms an emulsion with grease and fat. That is, it breaks the grease up into very small particles with which it intimately mixes. The soap dissolved in the water surrounding the globules prevents them uniting. In this condition they, and the particles of dirt with them, are easily washed away with the aid of much water.

Soaps are the sodium salts of the fatty acids, one of which is stearic acid. Consider a soap which is largely sodium stearate. Mixed with distilled water and shaken, a lather will be formed immediately; but when the soap is mixed with water containing dissolved calcium salts, the sodium stearate reacts with the calcium sulphate (for example) to form calcium stearate and sodium sulphate. If we let St stand for the stearate radicle, the equation representing the reaction is as follows:

$$2NaSt + CaSO_4 = CaSt_2 + Na_2SO_4.$$
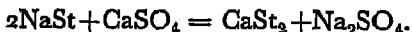
The calcium stearate is insoluble and appears as the scum which floats on the top of the water when soap is shaken with hard water. Not until all the calcium salts have been changed to calcium stearate can the soap form a permanent lather and so become a cleansing agent. Soap, then, is a water softener.

THE SOFTENING OF WATER

There are several methods possible, some of which apply to both kinds of hardness, temporary and permanent; others apply only to one kind.

1. Distillation

Obviously if water is distilled it must be soft as it contains no dissolved substances. The method is expensive but many laundries find it worth while to distil their water. Spent steam from turbines is used for this purpose. For some purposes the water must be not only soft, but free from certain, or all, dissolved substances. Then this method is essential.

2. Boiling

This removes the temporary hardness. The process is expensive and only partial, for some waters have more permanent hardness than temporary. Nevertheless it is an important method for small-scale use. The chemistry of the method is given on page 274.

3. Clark's process

This removes temporary hardness. The method is often employed to soften the water supplied to towns. The water to be softened is analysed and just sufficient lime added to react with the bicarbonate present:

$$Ca(HCO_3)_2 + Ca(OH)_2 = 2CaCO_3 + 2H_2O,$$
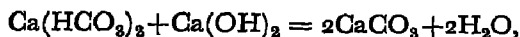
This causes the precipitation of calcium carbonate and so removes that part of the hardness which was due to the calcium bicarbonate, i.e. the temporary hardness. Hardness due to the presence of magnesium bicarbonate is also removed.

It is necessary in this process to add the correct amount. Any excess produces hardness owing to the calcium hydroxide in the water. A careless use of this method might result in the water being harder than before.

4. Adding washing soda

This removes both the permanent and temporary hardness. In the case of temporary hardness the calcium and magnesium bicarbonates, and in the case of permanent hardness the calcium and magnesium sulphates and chlorides, are converted into the corresponding sodium salts which do not cause hardness, e.g.

$$Ca(HCO_3)_2 + Na_2CO_3 = 2NaHCO_3 + CaCO_3,$$
$$CaSO_4 + Na_2CO_3 = Na_2SO_4 + CaCO_3.$$
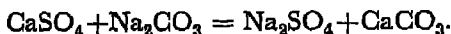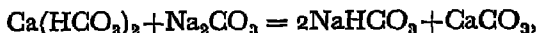
The calcium in both cases is precipitated as calcium carbonate. This method is used commonly in the home but too much washing soda is not good for the hands. This process is also sometimes combined with Clark's process in the softening of a town's water supply. The lime is added to remove the temporary hardness, and then sufficient washing soda is added to remove the permanent hardness. It is too expensive to use washing soda alone on a large scale.

5. The permutit process (see Fig. 125)

This is the modern method of softening the domestic supply of water. The water from a tap is passed into a cylinder and drawn off at the other end, softened; or the whole of the domestic supply may be passed through, so that all the water in the house is softened.

The substance which brings about the softening is called *permutit*. It is a complex substance having the appearance of sand. When water containing dissolved calcium salts is passed through this substance, the calcium salts are changed to sodium salts which do not cause hardness. If we let P represent the 'permutit radicle' then:

$$Na_2P + CaSO_4 = CaP + Na_2SO_4.$$
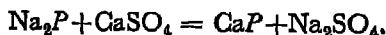
In time the permutit becomes exhausted because it has been deprived of its sodium, having taken calcium in exchange. The permutit can be regenerated by bringing it in contact with a solution of common salt, when the reverse process takes place, the sodium of the salt taking the place of the calcium in the

exhausted permutit. The process of softening water then can
be continued.

$$CaP + 2NaCl = CaCl_2 + Na_2P.$$
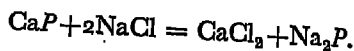
This process undoubtedly softens the water effectively; but
it is sometimes claimed that it removes all dissolved solids.
This is obviously wrong. It merely changes the calcium and
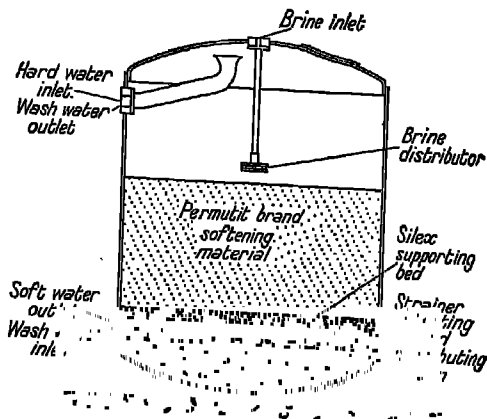
Fig. 125. Internal arrangements of permutit softener.

magnesium salts into the corresponding sodium salts. This,
of course, is all that is necessary to produce soft water. It does
not give pure water.

DOMESTIC WATER

We have dealt previously with the water supply of towns
(p. 13). It is essential for health that there should be a good
water supply in every house, for drinking, cooking, and cleaning.
All towns of any size in this country have such a supply, but
many country places too often have to depend on a supply not
above suspicion, e.g. unsuitable wells and streams.

Water from wells and streams may be quite good, but all
wells are not properly looked after, and some streams are
polluted. The most serious danger is that water should become

contaminated with organic matter, for in such water dangerous bacteria may be present. Water which contains organic nitrogen compounds is not safe.

It is clear that the water of streams passing through culti-vated areas cannot be fit for drinking because it is certain to be contaminated by the drainage from the fields. This does not mean that the water is necessarily dangerous, but that it might be so. Mountain and upland streams are nearly always free from serious contamination.

The water from shallow wells, especially those situated near farms, should always be suspected. Even the water of deep wells is not always safe, because there might be drainage into the well of surface water from cultivated fields. There is less chance of this happening if the inside of the well is bricked, and if the wall is also raised above the surface of the ground.

If you are suspicious of the water supply and you must drink the water, it should be first boiled. This kills all the bacteria, both harmful and harmless (p. 493).

Hard water for household use has obvious disadvantages. It is not so pleasant to wash with, and if it is very hard it is not pleasant to drink. Extremely hard water regularly used may be bad for health but the drinking of ordinary hard water supplied to towns in this country is not harmful to people in normal health.

Hard water also causes a hard deposit to form on the inside of kettles and domestic boilers. It consists of the carbonates of calcium and magnesium, and is formed during boiling from the bicarbonates of the calcium and magnesium present in the water. When the deposit (called 'fur') becomes thick much heat is wasted.

Water which is very soft is not pleasant to drink, and many people prefer ordinary hard tap-water to that softened by the permutit process. Soft water in the presence of air attacks lead. If water is conveyed by lead pipes in a building, there is some danger of lead poisoning if the water is too soft. Opinions differ on this question. Glasgow water is very soft, but no cases of lead poisoning through water have been reported. It is certainly safe for hard water to pass through lead pipes,

because the calcium sulphate in the hard water reacts with the lead to form lead sulphate. This remains as a hard insoluble coating on the inside of the pipe and preserves the lead from further action.

The waters supplied by different towns differ very much in hardness. This is to be expected as the supplies are drawn from different localities. Water from a granite district contains little dissolved matter and is soft; water from a limestone district is hard. It is only when water is too hard for domestic purposes that it is first treated before being distributed. There is no objection to a little hardness except, of course, the waste of soap it entails.

Tap-water is nearly always slightly alkaline. This can easily be tested with red litmus paper. This may be caused by a slight excess of lime added during Clark's softening process, or it may be due to the presence of calcium bicarbonate.

INDUSTRIAL WATER

Much water is required for industrial purposes, not only for raising steam but as a constituent in actual manufacturing processes. The hardness of the water is therefore important. An untreated hard water used in boilers soon causes heavy deposits which are expensive to remove. Hence it is desirable that boiler water should be soft. In addition it should not be corrosive, otherwise the metal of the boiler gradually disappears.

The supply of water in a district is an important consideration when a new site for a factory is being considered. Students of Geography claim that in many cases the development of a whole district depends on the hardness of the water available. It is said, for instance, that the hard water coming from the limestone hills dividing the north-east of Lancashire from Yorkshire has prevented the industrial development of North Lancashire, while the soft water of south-east Lancashire helped its industrial development.

If no suitable water for a particular industry can be obtained, the water must be treated before use. There is no general

method which can be applied to all industries. Each industry has its own particular needs, and each district its own particular water. The problem of the chemist in charge is to make the water suit the industry, and he does not always find it easy of solution. It is for this reason that the site of the works near a suitable water-supply is of great importance.

In softening the water for industry, four important considerations must be kept in mind:

1. The calcium and magnesium salts must be removed from solution.
2. The salts which come out of solution must not set into hard scales, but must be such as can easily be blown out.
3. The substances added to bring about this change must not cause any corrosion.
4. The process must be as cheap as possible.

WATER AS A SOLVENT

When we say that a substance is soluble in water we usually mean that it is *obviously* soluble. When we say that it is insoluble in water we mean that for all practical purposes, or for the purpose we have in mind, it is insoluble. Water will dissolve most substances to some slight extent, and those usually considered insoluble are not so in the strict sense of the word. Glass, for instance, is considered insoluble in water, and for most practical purposes is rightly so considered; but if water is kept in contact with glass for a long time, some of the substances in the glass dissolve to a very slight extent in the water. Water dissolves so many substances to a slight degree that great precautions have to be taken if it is desired to keep water as chemically pure as possible. Gold, platinum, and pure tin do not dissolve in water to any measurable extent.

Usually water will dissolve more of a solid with increase in temperature, but this is not always so. Slaked lime, for example, is more soluble in cold water than in hot. If a substance is very soluble in cold water, it does not necessarily mean that it will be very soluble in hot water. Common salt is only slightly more soluble in hot water than in cold.

When speaking about the solubility of a substance in water it is necessary to state the temperature. The *solubility of a substance in water* at a given temperature is the weight (in grams) of the substance which will dissolve *at that temperature* in 100 gm. of water.

. **Expt. 200. To find the solubility of potassium nitrate at various temperatures.** Dissolve some potassium nitrate in distilled water in a beaker. Warm the beaker over a bunsen flame, stirring the solution all the time with a glass rod. Keep a thermometer all the time in the solution. Do not allow the temperature to rise quickly but very slowly. When the desired temperature is reached and some solid remains undissolved, decant some of the solution into a weighed evaporating dish. Weigh the dish with the solution. Evaporate the solution, but not to dryness, otherwise the potassium nitrate will decompose. When the solution is concentrated, complete the evaporation on a water-bath. This will drive off all the water without allowing the temperature to rise higher than the boiling-point of water. When you think all the water has evaporated, dry the evaporating dish, and when cool weigh it with the potassium nitrate. Heat again for a few minutes on the water-bath and re-weigh after cooling. If this weight differs from the previous weight, repeat until the weight is constant, to ensure that all the water has been given off. Tabulate the results as follows:

	Grams
Weight of dish	
Weight of dish+solution	
Weight of dish+potassium nitrate: (i). . . .	
(ii) Constant . .	
From these three weighings find:	
Weight of solution taken	
Weight of potassium nitrate obtained	x
Weight of water evaporated	y

Therefore 100 gm. of water dissolve $100x/y$ gm. of potassium nitrate. This is its solubility at that particular temperature. As 1 c.c. of water weighs 1 gm., this is also the weight that will dissolve in 100 c.c. of water.

In this way the solubility of the salt at other temperatures can be found. A convenient way of showing the results is to draw a graph. Fig. 126 shows such a graph drawn for various salts.

It is often important to know whether a substance is soluble or insoluble in water. All potassium, sodium, and ammonium salts are soluble in water; and so are all nitrates, nitrites, and

chlorates. All common chlorides, with the exception of three, are soluble in water. The three exceptions are lead chloride, which is sparingly soluble in cold water but readily soluble in hot water, mercurous chloride, and silver chloride. The solubility of slaked lime (calcium hydroxide) at 0° C. is twice as great as its solubility at 100° C. On the other hand,
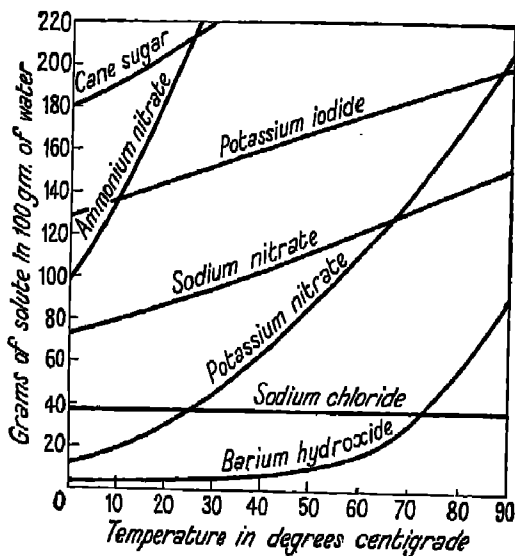
FIG. 126. Solubility curves of some common substances.

the solubility of potassium chlorate is doubled as the temperature rises from 0° to 100° C. Common salt is more soluble than potassium chloride at 0° C., but at about 20° C. their solubilities are equal, while at 100° C. potassium chloride is much more soluble than common salt. Sodium nitrate is much more soluble than potassium nitrate at 0° C., but at 100° C. the position is reversed. The comparative solubility of salts is best studied from graphs. Refer to Fig. 126 and try to read from the graph other facts similar to the above.

PROPERTIES OF WATER. TESTS FOR WATER

As we have seen previously, a useful test for water is to add it to anhydrous copper sulphate which turns blue. This, of course, is not a test for the purity of the water. It merely shows that some water is present.

Pure water is a colourless, odourless liquid, neutral to litmus, and leaving no residue on evaporation. It boils at 100° C. (at 760 mm. pressure) and freezes at 0° C. Its density is 1 gm. per c.c. at 4° C.

A liquid having all these properties is not necessarily pure water, but it would at least be almost chemically pure if the tests were properly carried out. It might be very 'impure' as a drinking water, for it could contain harmful bacteria, and still pass all the above tests.

QUESTIONS

1. What is meant when it is said that water is hard? How does temporary hardness differ from permanent hardness?

2. State and explain, giving equations, the methods of softening water.

3. What are the causes of the hardness of water? Which natural waters do you expect to be hard and which soft? Give reasons for your answer.

4. Two samples of water A and B gave the following results when treated with soap solution: 50 c.c. of A, unboiled, required 6 c.c. of soap solution, boiled, 1 c.c.; 50 c.c. of B, unboiled, required 10 c.c. of soap solution, boiled, 8 c.c. Interpret these results.

5. What is soap and how is it made? What happens when soap is used with hard water? Why is soap a good cleansing agent?

6. What kind of water is desirable for domestic purposes? What are the usual sources of domestic water and what precautions must be taken in selecting water for this purpose?

7. What water problems confront an industry beginning in a new district and how are they dealt with?

8. What are the properties of pure water? How would you show that a liquid contained water? How can water be tested for purity?

XVIII

SOME ORGANIC SUBSTANCES

THE COMPOSITION OF LIVING MATTER

THE greater part of the dry weight of living matter consists of three classes of substances: Carbohydrates, Proteins, Fats. There are many other substances in living matter not belonging to these classes, but the bulk of animals and plants, apart from water, is made up of these three substances.

These substances also form the food of animals and plants. We may now use the chemistry we have learned to understand more about these important substances. Some of them are very complex in structure, so complex indeed that their complete structure is unknown, but much already is known about them, and our knowledge is increasing almost day by day. We shall proceed to study each class in turn.

CARBOHYDRATES

The chief carbohydrates occurring in plants are starches, sugars, and cellulose. There is a very important carbohydrate which occurs in animals, chiefly in the liver. It is called *glycogen*, and is sometimes referred to as 'animal starch'.

The commonest carbohydrate occurring in plants is cellulose, as this substance forms the walls of the cells. Some walls are subsequently changed, but all begin as, and a great number remain as, cellulose.

Starch and sugar are found somewhere in most green plants, some plants being richer in one than the other, while many monocotyledons, e.g. onion, contain only sugar. The root of the sugar beet is specially rich in sugar, while the potato tuber is rich in starch (p. 135).

All carbohydrates are compounds of carbon, hydrogen, and oxygen, the last two elements being present in the same proportions as in water. A general formula for carbohydrates can be written as $C_xH_{2y}O_y$ where x and y represent (usually)

different numbers. We could write the formula $C_x(H_2O)_y$ and this would remind us that the elements hydrogen and oxygen are present in the same proportion as in water. It is better, however, to use the first formula, for the second might lead us to think that a carbohydrate is a compound of carbon and water. We have seen (p. 137) that this is not so.

STARCH

We have already learned something about this compound (p. 135). The commercial form is obtained from potatoes, rice, wheat, or maize. It is insoluble in cold water, but on boiling the granules burst and form a nearly clear solution which sets to a jelly on cooling. Its simplest formula is $C_6H_{10}O_5$, but it is usually written $(C_6H_{10}O_5)_x$ to indicate that the size of the molecule is unknown, but that the number and kind of atoms in the molecule are in the proportion shown by that part of the formula in the brackets. When boiled with acids it is changed to a sugar which has the formula $C_6H_{12}O_6$.

As we have already seen, the acid appears to act as a catalyst. The starch reacts with water to form glucose. This is the final result, but there are intermediate stages in the reaction.

The changing of starch to sugar may be brought about without the action of acids. This is done in the plant by means of an enzyme called *diastase* (also called *amylase*).

THE ACTION OF DIASTASE ON STARCH

Expt. 201. Put a little starch paste in a test-tube of water and shake. Divide the starch 'solution' thus made into three equal portions. To one add iodine solution. Note the blue colour indicating the presence of starch. The solution must be *cold*, as the blue colour disappears on warming. (Try this.) To the second portion add a pinch of diastase. (This can be purchased.) Shake the test-tube to mix the contents and place it in a beaker of water which is kept at a temperature of about 40° C. To the third portion add nothing, but put the test-tube in the water-bath with the second test-tube. Allow them to stand for five minutes in the bath.

Put a drop of the second portion into a watch-glass containing a little dilute iodine solution. Note the colour change, if any. Repeat every five minutes until there is no blue colour, showing that all the starch has disappeared. Test the remainder of the second portion by boiling with Fehling's solution. Sugar is present. Test the third

portion, which has been in the water-bath all the time, for sugar. None is present. Test for starch. It is still present.

The diastase converts the starch to a sugar called maltose, or malt sugar ($C_{12}H_{22}O_{11}$). This sugar when heated with Fehling's solution causes a reddish-brown precipitate. In this respect it is like glucose, or grape sugar ($C_6H_{12}O_6$). Sugars which do this are called reducing sugars.

ENZYMES

Diastase belongs to a class of substances called *enzymes*. They were formerly known as ferments, and this term is still frequently used. Enzymes are probably proteins and are formed by living protoplasm, but they are not themselves living bodies. They are destroyed by heat above about 60° C., and work best between 25°–40° C. They are very much like inorganic catalysts. They cause the chemical change to take place without appearing to enter into the reaction themselves. Unlike inorganic catalysts, however, they are themselves destroyed to a small extent in the process.

The enzyme diastase occurs in barley seeds, in the leaves of plants, and in animal bodies. It is always concerned with the same chemical reaction or its reverse according to conditions, i.e. it either helps starch to react with water to form maltose or it helps to build up starch from sugar.

There are other enzymes which can promote other chemical reactions. The enzyme maltase can change maltose into glucose.

$$C_{12}H_{22}O_{11} + H_2O = 2C_6H_{12}O_6.$$

We shall meet other enzymes in the course of our study, but all of them are capable of performing one operation only, or one closely allied to it. In this again they differ from inorganic catalysts some of which, e.g. finely divided platinum, can promote different reactions.

THE ACTION OF SALIVA ON STARCH

Expt. 202. Collect a little of your saliva in a test-tube and dilute it with thirty times its volume of distilled water. Divide the liquid into six portions (*A* to *F*) and test as follows:

A. Add Fehling's solution and show that it contains no sugar.

B. Test with litmus paper and show that it is alkaline. (If it is acid you should start again and take the specimen of saliva after you have washed your mouth out with water.)

C. To 10 c.c. add 2 c.c. of 1 per cent. starch solution and keep it at 40° C. in a water bath. Withdraw a drop at intervals of two minutes and test with iodine. Continue this until the last test shows that starch is absent. Test the remainder with Fehling's solution and show that sugar is present.

D. To the saliva add starch solution and a drop of sodium carbonate solution, and treat as *C*.

E. To the saliva add starch solution and a drop of dilute hydrochloric acid until faintly acid in reaction, and treat as *C*.

F. Boil the saliva, cool it, then add starch solution and treat as *C*.

G. As a control experiment treat as *C* but add distilled water instead of saliva.

In human saliva there is an enzyme called *ptyalin*, which is very much like diastase in that it changes starch to maltose. Saliva is slightly alkaline in reaction. Ptyalin appears to be active in an alkaline, neutral, or fairly acid medium, but it is inactive in a stronger acid solution. Its particular chemical action is to convert starch to sugar (maltose). It has no action on proteins or fats. Another enzyme changes the maltose into glucose.

THE SUGARS

Sugars are all soluble in water and have a sweet taste, but some are much sweeter than others.

Expt. 203. Divide a solution of ordinary table sugar (cane sugar or sucrose) into three portions. Test the first portion for starch with iodine solution. Test the second portion with Fehling's solution. Boil the third portion with a little dilute hydrochloric or sulphuric acid, neutralize the acid by adding caustic soda solution, and test with Fehling's solution.

You will find no starch in the first portion, and there will be no precipitate with Fehling's solution in the second portion.

With the third portion, however, there will be a precipitate. Glucose is present. When sucrose is boiled with a dilute acid each molecule of sucrose reacts with one molecule of water to form one molecule of glucose and one molecule of fructose. This action can also be brought about by means of the enzyme sucrase (also called invertase).

$$C_{12}H_{22}O_{11} + H_2O = C_6H_{12}O_6 + C_6H_{12}O_6.$$
Sucrose   Glucose   Fructose
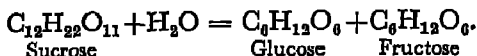
Fructose is a sugar very much like glucose and has the same formula. It differs in the arrangement of the atoms in the molecule but not in the number of atoms in the molecule.

In this reaction one molecule of a substance by reacting with water becomes two molecules. A chemical action in which a compound is split up into two or more compounds by interaction with water is termed *hydrolysis*. This word by derivation means 'splitting up by means of water' just as *electrolysis* literally means 'splitting up by means of electricity'. We can say that the sucrose is hydrolysed to glucose and fructose.

The common sugars may be divided into two classes, the simple sugars, which are not hydrolysed in acid solutions, and the compound sugars which are hydrolysed in acid solutions or by enzymes. The simple sugars have the general formula $C_6H_{12}O_6$. Important examples are glucose, sometimes called dextrose or grape sugar, and fructose, sometimes called laevulose or fruit sugar.

The commonest compound sugars have the general formula $C_{14}H_{22}O_{11}$. The most important example of this class is sucrose or cane sugar. Two other important ones are maltose, or malt sugar, and lactose, or milk sugar. The last-named occurs in milk.

All the sugars mentioned in both classes, with the important exception of cane sugar, give the reddish-brown precipitate on boiling with Fehling's solution. They are REDUCING SUGARS. All the compound sugars are hydrolysed to simple sugars when heated with acids. They are also hydrolysed by enzymes.

On hydrolysis:

$$\left.\begin{array}{l} \text{1 mol. of cane sugar} \\ \text{1 mol. of maltose} \\ \text{1 mol. of lactose} \end{array}\right\} \rightarrow \text{1 mol. of glucose} + \left\{\begin{array}{l} \text{1 mol. of fructose} \\ \text{1 mol. of glucose} \\ \text{1 mol. of galactose} \end{array}\right.$$

$$C_{12}H_{22}O_{11} + H_2O = C_6H_{12}O_6 + C_6H_{12}O_6.$$
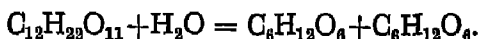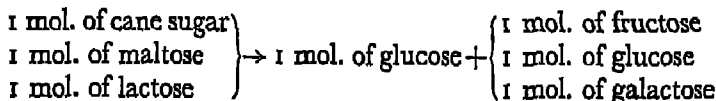
In each case at least one molecule of glucose is produced.

Expt. 204. Mash with water some white beetroot and boil for a few minutes. Filter, and cool the filtrate. Divide it into three portions and treat each as detailed in Experiment 203. Deal in the same way with such materials as carrot, cabbage leaves, potato, flour, &c., and try to discover if starch, cane sugar, and the reducing sugars are present.

CELLULOSE $(C_6H_{10}O_5)_x$

This carbohydrate, which has the same general formula as starch, forms the cell walls of plants. Linen, made from the fibres of the stems of the flax plant, and cotton, the hairs occurring with the cotton seed, are examples of almost pure cellulose. Cotton wool is largely cellulose. Paper, which is made from linen, cotton, and wood, is also largely cellulose. Filter paper is practically pure cellulose. Gun-cotton, from which cordite (a smokeless explosive) is made, collodion, used in the manufacture of photographic films, and celluloid are all produced from cellulose. Artificial silk, cellophane, and 'windowlite' (cellulose acetate) are all made from wood pulp.

Cellulose can be hydrolysed to sugar by the action of acids, but the human body contains no enzyme that can hydrolyse cellulose. Starch taken into the human body is hydrolysed to glucose, but cellulose remains un-hydrolysed, and is passed out of the body unaltered. Herbivorous animals, however, are able to digest cellulose to some extent.

A CLASSIFICATION OF SOME CARBOHYDRATES

General formula $C_xH_{2y}O_y$

I. SOLUBLE IN WATER.

   A. Simple sugars. General formula $C_6H_{12}O_6$. Not hydrolysed,

      e.g. Glucose $\Big\}$ Reducing sugars.
            Fructose

   B. Compound Sugars. General formula $C_{12}H_{22}O_{11}$. Hydrolysed in acid solutions and by enzymes.

      e.g. Cane sugar (Sucrose)—Non-reducing.

           Malt sugar (Maltose)—Reducing.

           Milk sugar (Lactose)—Reducing.

II. INSOLUBLE IN WATER. General formula $(C_6H_{10}O_5)_x$. Hydrolysed in acid solutions and by enzymes:

      e.g. Cellulose.

          Starch.

          Glycogen ('Animal Starch').

HYDROLYSIS OF FOOD RESERVES IN PLANTS

Expt. 205. Place some seeds (barley will do very well) on a piece of muslin stretched over a beaker of water. Allow some of the muslin to trail in the water so that the seeds have a constant supply of water without being deprived of air. Place the whole in a warm, but not hot, place. Place some more of the same kind of seeds under the same conditions except that they must be deprived of water. Leave for a few days. The damp seeds will sprout in time, but the dry seeds will not do so.

The seeds contain a reserve supply of starch as food (p.140). (Crush some of the dry seeds and test for starch.) This starch is for the use of the young plant until it has produced green leaves and can manufacture food for itself. The young plant (called the embryo) which is already present in the dry seed cannot use the insoluble starch. It must first be changed into some soluble substance. Diastase is present in the seed and can change the starch into soluble sugar, maltose. This change, however, cannot begin until water is present because the elements of water are required in the chemical action. This is one

reason why seeds cannot germinate unless there is an adequate supply of water. The water is required for the hydrolysis of the starch. It is also required to act as the solvent for the products of the hydrolysis.

The hydrolysis of the food in a seed is only a particular case of the hydrolysis which must go on in all food reserves where the food is unsuitable for assimilation.

Many plants, called *biennials*, spend the first year of their life in storing food to be used in the second year to promote a quick and vigorous start towards the ultimate production of seeds, by which it can reproduce itself in great numbers. The *turnip* is such a plant; but man takes the plant after its first year's growth, and uses the stored food for himself. When, however, he requires turnip seeds, he allows the plant to continue for another season, during which the food reserve in the root disappears and the plant produces seeds. The plant manufactures food all the time for itself. The reserve supply is used by the plant at a time when growth is so vigorous that it could not supply itself with all the food necessary in the time. It spends a year storing, and then a year using the stored food as an extra supply.

Examine the root of a turnip that has 'gone to seed'. Compare it with a turnip root at the end of its first year. The one is flabby and stringy, while the other is firm and hard.

The cultivated *cabbage* does not normally go to seed the first season, but stores food in the leaves which become fleshy. If, however, the cabbage is not pulled, but the fleshy leaves cut at the top to allow the stem to develop, the plant produces seed and at the same time the leaves lose the stored food and become flabby instead of being crisp.

Sometimes cabbages, and other plants of the same family, do go to seed in the first year. This is called 'bolting' by the gardener who does not want the seeds but does require the plant for food.

A *potato* is the swollen end of an underground stem (a tuber). It is a reserve food store produced by the plant for the purpose of reproduction without seeds (Ch. XLVIII). In the cultivated state the tubers are pulled up and used by man, mostly for food,

but some for the production of more potato plants. A potato plant left unpulled dies down in the autumn, but in the following spring new plants arise from the tubers produced the previous year. Naturally the plants are crowded, and some die because of overcrowding and the consequent insufficiency of water and food in the soil. 'Wild' potatoes are always small and of poor quality.

Put some good, firm potato tubers on a dry tray and leave them in the dark. Watch them from day to day. They will begin to sprout from the 'eyes'. Note the flabbiness of the sprouting potato compared with the firmness of the unsprouted tuber. Test a little of the potato for sugar with Fehling's solution. You will find sugar present in the sprouting tuber, but little or none present in the unsprouted tuber.

Dig up a potato plant that is growing vigorously. Note the small tubers already formed and try to find the shrunken old tuber from which the plant started as a small shoot.

The starch in the old tuber is hydrolysed to sugar. This is done by means of the enzyme diastase. The sugar is then transported to the growing parts of a plant and is used for the production of more protoplasm, and expended in the production of energy. The water necessary for the hydrolysis is taken from the tuber itself until roots have been formed, when water can be drawn from the soil. As the starch disappears the old tuber decreases in size until there is little more left than a shrunken skin. By this time, however, the plant is in a position to manufacture food for itself.

BREWING

The object in brewing is to produce a beer which shall be palatable and shall contain from 2 to 6 per cent. of alcohol.

Barley grain is spread on a malting floor in such conditions of moisture and temperature that germination is brought about. When the enzyme diastase, which is present in the grain, has caused the starch of the grain to react with water to form the sugar maltose, the young plants are killed by raising the temperature. This is done so that the maltose will not be used as food by the young plants.

The maltose is then extracted by dissolving it in hot water. Hops are added to the solution to give the beer the flavour

required, and then special yeast cultures are added according
to the type of beer required.

Yeast is a fungus of microscopic size.

Put a very little yeast in water on a slide and examine it under the
high power of the microscope (Fig. 127).

The complete plant is a very tiny cell devoid of chlorophyll
and so cannot manufacture sugar from the carbon dioxide of
the air. Yeast must therefore obtain its food in another way.
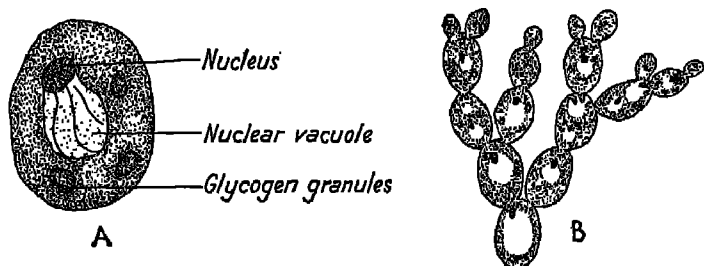
FIG. 127. Yeast.
A. Single individual. ×2500. *B.* Multiplication by budding.

At a temperature of about 25° C. the yeast grows quickly in
the solution obtained from the germinated barley. As it grows
it reproduces itself by budding. The sugar gradually disappears,
some of it being used by the yeast as food to enable it to grow,
and some of it being decomposed by the yeast to supply the
energy required for its growth. The result of the decomposition
is that in the main carbon dioxide and alcohol are formed.

The maltose is first hydrolysed by the maltase in the yeast to
glucose. This is then acted on by another enzyme in the yeast
called *zymase*. This decomposes the glucose into carbon
dioxide and alcohol.

$$C_6H_{12}O_6 = 2C_2H_5.OH + 2CO_2.$$

The process is called *Alcoholic Fermentation.*

This is an example of an action brought about by an enzyme
but which is *not* hydrolysis. No water enters into the action.

The enzyme, *invertase*, is also present in the yeast and this hydrolyses cane sugar, when present, into glucose and fructose.

Grape juice is fermentable without the addition of yeast because there is a yeast always present in the soil in which the vine grows. This yeast finds its way on to the grapes from the soil. By the fermentation of the juice of grapes wines are made.

Spirits normally contain 40–50 per cent. of alcohol. They are made by the distillation of fermented alcoholic liquid. Whisky is made by distilling the 'wash' from the action of yeast on malted barley or rye. Rum is made by fermenting molasses with yeast and distilling the product.

Expt. 206. Fermentation of sugar solution. Place about 500 c.c. of a 5 per cent. sugar solution in a large flask and add some brewer's yeast. Arrange a delivery tube from the flask to pass into lime-water. Keep in a warm place (about 25° C.) for several days. The lime-water will turn cloudy showing the evolution of carbon dioxide, and the contents of the flask will be frothing and will smell of alcohol.

Place a little of the liquid on a slide and examine it under the high power of the microscope. The yeast cells (Fig. 127) will show signs of budding showing that growth has taken place.

Distil off the liquid in the flask and collect the fraction which distils below 90° C. This fraction should be redistilled and the fraction distilling below 80° C. collected. This can be shown to be alcohol by smelling and by burning. Use the apparatus shown in Fig. 4, but insert a thermometer through the cork so that the bulb is opposite to the exit.

ALCOHOL

The alcohol produced in the previous experiment is only one of many organic compounds classified as alcohols. This particular one is known as ethyl alcohol. By very careful distillation most of the water can be removed, and a 96 per cent. alcohol obtained. It is difficult to remove the remaining water, but it can be done, and the alcohol is then known as *absolute alcohol*. This liquid is colourless, has a distinct smell and taste, and burns in air to form carbon dioxide and water.

$$C_2H_5.OH + 3O_2 = 3H_2O + 2CO_2$$
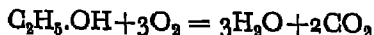
Apart from its use in the form of alcoholic beverages, it has

a great industrial demand. Many industries find a use for it.
It is employed as a solvent for varnishes, lacquers, enamel,
perfumes, &c. The anaesthetics chloroform and ether, and the
disinfectant iodoform are manufactured from it, and mixed
with benzine it is used to a limited extent as a motor fuel. If
alcohol and a little ether are mixed with collodion the so-
called 'solid alcohol' is obtained. This jelly burns away without
leaving any solid or liquid residue.

Methylated spirit is alcohol mixed with other substances
giving it an unpleasant smell and taste, thus making it most
unpalatable. There is no duty on methylated spirit so it is
comparatively cheap. On the other hand, there is a heavy tax
on alcohol, consequently it is rather expensive.

Beer or wine exposed to the air turns sour. This is due to the
presence of certain bacteria (*mycoderma*) causing the oxidation
of the alcohol by enzyme action. This is the basis of the pro-
duction of *vinegar* and *acetic acid*.

PROTEINS

We must now study another class of substances which form
a large part of living matter. Indeed there is no living matter
without proteins. We do not know what living protoplasm is,
but dead protoplasm seems to consist largely of very complex
proteins. Even the simplest proteins are so complex that the
structure of the molecule is not known. Nevertheless much is
known about them, much more than we shall require to know
in the present course.

Proteins can be broken down to substances called *amino-acids*.
These are complex substances containing nitrogen, but they
are simple compared with proteins. Protein molecules consist
of complicated chains of amino-acids. The amino-acids are
the units of which protein molecules are built, somewhat as
bricks are the units of which a house is built. The chemist can
break down proteins to amino-acids but he cannot build up
proteins from amino-acids. Both these processes take place in
an animal's body. The molecules of proteins are so large that
some students of the subject think that the term molecule should
not be applied to the ultimate particle of a protein. The large

size of the ultimate particles of proteins has important effects which we must study.

Flesh without fat is almost pure protein, and so is the white of an egg. There is a very large number of proteins and they differ very much in composition and in properties, but they all consist of carbon, hydrogen, oxygen, nitrogen, and sometimes sulphur and phosphorus. Some, like the white of an egg (egg-albumen), are soluble in water. Others, like fibrin (a protein of the blood), are insoluble in water, while some are soluble only in strong acids and alkalis.

SOME PROTEIN REACTIONS

Expt. 207. Examination of egg-albumen. Examine some of the white of a raw egg, which is almost entirely a solution of a protein called egg-albumen. It is slimy, viscous, and non-crystalline. Put a little of the egg-albumen in water and warm gently. It coagulates, changing into an opaque white mass. Heat a little undiluted egg-albumen in a dry test-tube. Test the vapour which condenses on the side of the tube for water. On further heating acrid fumes are given off. Test the fumes with red litmus paper and with a glass rod which has been dipped in hydrochloric acid. It will be found that the fumes contain ammonia, thus indicating the presence of nitrogen in the protein. The fumes will also be found to be inflammable, burning with the smoky flame of hydrocarbons. A black deposit of carbon is left in the tube. We have therefore demonstrated the presence of carbon and nitrogen. We also know that hydrogen and oxygen were also present as water was given off.

The presence of nitrogen in protein matter can also be demonstrated by heating it with soda-lime, ammonia being evolved.

Beat up some more egg-albumen with about ten times its volume of water. The egg-albumen dissolves. Use this solution for the following tests:

Expt. 208. Biuret test. To a little of the albumen solution add about half the volume of dilute caustic soda solution. To this add 1 per cent. copper sulphate solution, drop by drop, until a violet colour is obtained. The colour may be so faint that it cannot be seen unless viewed along the length of the test-tube.

This test applies to all proteins and not merely to egg-albumen. Sometimes the colour is pink and not violet.

Expt. 209. The xanthoproteic reaction. To a little of the albumen solution add about half the volume of strong nitric acid. Boil. The solution turns yellow. Neutralize with ammonia solution. The colour changes to orange.

This reaction occurs with most proteins but not with all. Nitric acid stains the fingers yellow. This is the same action.

Expt. 210. Millon's reaction. To a solution of the albumen add about half its volume of Millon's reagent (a mixture of mercuric nitrate and mercurous nitrate obtained by dissolving mercury in a suitable quantity of nitric acid). A white precipitate is obtained which turns to brick red on heating. If the protein solution is weak there may be no precipitate but the solution turns brick red on boiling.

Expt. 211. To test for sulphur in egg-albumen. Boil a little of the undiluted egg white for a minute with about half its volume of 40 per cent. caustic soda solution. Add a few drops of lead acetate solution. A black or brown precipitate of lead sulphide indicates that sulphur was present in the protein.

DIFFUSION OF CERTAIN SUBSTANCES THROUGH A MEMBRANE OF PARCHMENT

The passage of substances in solution through membranes is very important in the life of an animal or plant. Let us see what kind of substance will pass through membranes.

Expt. 212. To show that:

(i) A solution of common salt will pass through a membrane of parchment.

(ii) A solution of glucose will pass through a similar membrane.

(iii) A solution of the white of an egg (a protein) will not pass through a similar membrane.

(iv) A solution of starch will not pass through a similar membrane.

Prepare (or obtain) four bags of ordinary parchment paper like that we used in the osmosis experiments on p. 26. In (a) put a solution of common salt in distilled water, in (b) a solution of glucose, in (c) a solution of the white of an egg, and in (d) a solution of starch. Suspend the bags in separate beakers of distilled water and leave them there for a few hours. Test the water in which (a) is suspended for salt by adding silver nitrate solution. (A white precipitate indicates the presence of a chloride.) Test the water in (b) for glucose by Fehling's solution. Use the tests given above to see if there is any protein in the water in (c). Add iodine solution to the water in (d) to test for starch.

It will be found that salt and glucose diffuse through the membrane and that starch and egg-albumen do not.

COLLOIDS

Solutions of substances like salt and sugar diffuse through a membrane of parchment, but solutions of such substances as starch, gelatine, glue, and the white of an egg do not pass through such a membrane. The former are called *crystalloids* because most of the substances which diffuse quickly through a membrane of parchment are crystalline. Those substances which will not diffuse through a similar membrane are called *colloids*.

Starch and proteins form colloidal solutions. The molecules which exist in the solutions are too large to pass through the pores of the parchment. The molecules of salt and sugar, however, are much smaller and easily pass between the pores.

A colloidal solution is a suspension of extremely small particles, much too small to be seen under the most powerful microscope, and yet very large when compared with the size of the molecules of such substances as salt and sugar.

Most of the soluble part of our food is colloidal and therefore cannot pass through the membranes of the intestines into the blood. We shall see how the colloids are changed into crystalloids so that they can pass through the membranes of the intestines. In the case of starch it is hydrolysed to sugar by an enzyme.

THE ACTION OF PEPSIN ON PROTEIN

Let us now see how a protein is dealt with in digestion. Pepsin is an enzyme which promotes the change of proteins to simpler substances called peptones and these can be changed to amino-acids. Let us first try to discover how pepsin acts on protein.

Expt. 213. Take five test-tubes (*A–E*) and in each place a small cube of boiled white of egg. Also to each add a small drop of chloroform to prevent the action of bacteria. Add to:

(*A*) 2 c.c. of distilled water. This is the *control* experiment.
(*B*) 1 c.c. of pepsin solution and 1 c.c. of dilute hydrochloric acid.
(*C*) 1 c.c. of pepsin solution and 1 c.c. of dilute sodium carbonate solution so that it is slightly alkaline.
(*D*) 1 c.c. of boiled and then cooled pepsin solution and 1 c.c. of dilute hydrochloric acid.

(*E*) 1 c.c. of distilled water and 1 c.c. of dilute hydrochloric acid
solution to show that any action is not due to the action of the
hydrochloric acid alone.

Keep the tubes in a water-bath at 40° C. for two or three hours.
Note whether any of the protein has disappeared. Test the solution
with the Biuret test to see whether any protein has gone into solution.

It will be seen by these experiments that pepsin will only act upon
the protein in an acid medium. Heating destroys the action of the
enzyme.

Pepsin may be purchased in powder form. To make a solution
take 0·5 gm. of the powder and dissolve it in 100 c.c. of hydrochloric
acid containing 3·65 gm. of acid per litre of solution.

Pepsin is an enzyme which occurs in the stomach. In the
presence of hydrochloric acid, which also occurs in the stomach,
it has the power of changing proteins to simpler substances
called *peptones*. These peptones are soluble, and will diffuse
through a membrane of parchment but only very slowly.
Another enzyme in the intestines changes the peptones into
amino-acids before they are absorbed. The digestion of the
proteins, however, begins in the stomach. Pepsin has no effect
on starch. Test this statement.

FATS

Animal fats, such as suet, tallow, lard, and butter, occur in
the bodies of animals, and vegetable oils, such as palm, olive,
and castor oils, are stored in the seeds or fruit coats of many
plants. Although quite different in appearance these fats and
oils are very similar chemically. They are mainly mixtures of
compounds formed from glycerol (commonly called glycerine)
and certain acids. The chief acids concerned are two 'fatty'
acids—stearic acid (from Gr. for suet) and palmitic acid (from
'palm')—and another acid, oleic acid (from 'oil'). The com-
pounds resemble salts in their chemical composition and are
called *esters*. Glyceryl stearate and palmitate are solid but
glyceryl oleate is a liquid so that the relative amounts of the
three esters decide whether the mixture is hard, like suet, or
soft, like lard, or liquid, like olive oil. Another 'fatty' acid,
butyric acid, gets its name from butter.

We have seen that when a fat is boiled with caustic soda,
sodium salts of the acids are produced and glycerol set free.

The sodium and potassium salts of stearic, palmitic, and oleic acids are called soaps.

THE ACTION OF LIPASE ON FATS

Lipase is an enzyme which is secreted by the pancreas and passed into the small intestine (Ch.XLIV). It has the power of hydrolysing fats into glycerol and fatty acids. Fats and oils do not dissolve in water or in the liquids of the stomach and intestines, therefore the action of the enzyme is helped if the fat is in a very fine state of division so that a great surface of the fat is exposed to the action of the enzyme. This is the case when the fat is emulsified.

The fats are emulsified in the small intestine by an alkaline fluid, the bile, and lipase then acts on the fats.

Expt. 214. The action of lipase on olive oil. Take two test-tubes. In the first shake up some olive oil and water and in the second shake up some olive oil, water, and lipase. Allow both tubes to stand and compare the contents. In the second tube the oil and water mix and the oil is no longer visible, whereas in the first tube the liquid separates into two layers. Therefore lipase has acted on the fat.

SUMMARY

$$\text{Glycerol} + \begin{Bmatrix} \text{Stearic acid} \\ \text{Palmitic acid} \\ \text{Oleic acid} \end{Bmatrix} \begin{matrix} \text{Formation} \\ \text{of fat} \\ \xrightarrow{} \\ \xleftarrow{} \\ \text{Hydrolysis} \\ \text{by lipase} \end{matrix} \text{Glyceryl} \begin{Bmatrix} \text{stearate} \\ \text{palmitate} \\ \text{oleate} \end{Bmatrix} + \text{water}$$

$$\text{Glyceryl stearate, \&c.} + \begin{matrix} \text{Caustic soda} \rightarrow \\ \\ \text{Caustic potash} \rightarrow \end{matrix} \begin{Bmatrix} \text{Sodium stearate,} \\ \text{\&c.} \\ \text{Hard soap} \\ \text{Potassium stearate,} \\ \text{\&c.} \\ \text{Soft soap} \end{Bmatrix} + \text{Glycerol}$$
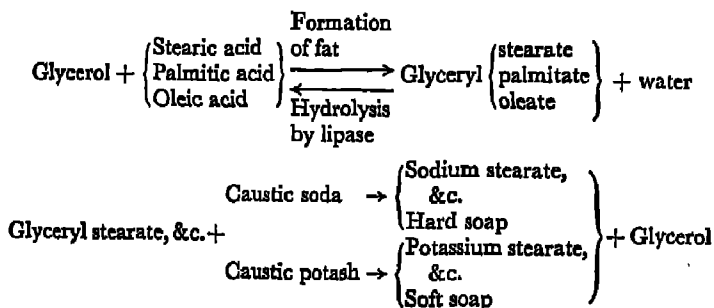
Use of soap in washing, i.e. making hard water soft:

Sodium stearate, &c. + calcium chloride
= Calcium stearate, &c. + sodium chloride.

QUESTIONS

1. Distinguish between carbohydrates, fats, and proteins. Name some common (*a*) carbohydrates, (*b*) fats which occur in (i) plants and (ii) animals.

2. How does cane sugar differ from glucose? What test would you make to discover which of these two sugars was present in a given solution? How may cane sugar be converted to glucose in the laboratory?

3. You are given a liquid containing both starch and glucose. Describe in detail how you would obtain the starch free from the glucose.

4. What is an enzyme? Give three examples and describe an experiment to illustrate the action of any one of them. How does an enzyme differ from an ordinary inorganic catalyst?

5. What is cellulose? Where does it occur and what are its uses?

6. What do you understand by the term hydrolysis? Illustrate your answer by giving two examples. How is this process concerned with the germination of a seed?

7. Describe the process of brewing beer, paying special attention to the uses made of the activities of barley and yeast.

8. Describe a laboratory experiment to illustrate the action of pepsin on protein. How would you test a protein for the presence of sulphur?

9. Of what does animal fat consist? What is an ester? Describe the action of caustic soda on fats.

XIX

FORCE

IF you try to raise a bucket of coal from the floor to a bench you will feel the effect of the effort you make in your muscles. Perhaps you will find the bucket too heavy to lift and you may seek the help of a friend. The two of you, lifting together, may manage to raise the bucket. Your effort alone did not raise it, but perhaps you felt that if you were a little stronger you could manage it. In other words, the effort (or force) you applied did not raise the bucket but tended to raise it. A force may be insufficient to raise a body but it may tend to raise it.

Suppose you wish to stop a light truck which is running downhill. You could get in front of it and so stop it; or you could pull it from behind. In either case you must apply a force to stop it. A force must be applied to stop a moving object.

A force is defined as *that which changes or tends to change the state of rest or motion of a body*. If a ball is moving downhill in a straight line and keeping the same speed (this is possible by special arrangement) you would have to apply a force to (1) stop it, (2) slow it down, (3) make it move faster, or (4) alter its direction.

A body may be pushed or pulled by another body. In both cases the bodies must be in contact. If you push a truck, your body must be in contact with it either directly or through another body such as a rod. There is a force between your body and the truck preventing (or tending to prevent) your body from penetrating that of the truck, and the truck from penetrating yours. Such a force is known as a *thrust*. When you stand on the floor there is a thrust downwards due to your weight and a thrust upwards preventing you from going through the floor.

If you pull the truck there is a force between you and the truck preventing (or tending to prevent) you and the truck from coming apart. This kind of force is called a *tension*. There

is a tension in the connecting chain of a railway truck and the
engine which is pulling it.

When a magnet is placed near a piece of iron the latter
moves (or tends to move) towards the magnet although the
two bodies are not in contact. We say that the magnet attracts
the iron. The force between the magnet and the iron is called
an *attraction*. Similarly, two electrified bodies may be made to
repel one another although the bodies are not in contact. The
force acting between the two is then called a *repulsion* (Ch.
XXXVI).

If you push against a wall (without going through it) there
is a thrust back. We may consider that there are two forces
acting, your own, called the *action*, and the push back from the
wall, called the *reaction*. Action and reaction are always equal
and opposite. Clearly it is so in this case. If your thrust were
greater than that of the wall, you would go through, while if
the thrust from the wall were greater than yours, you would be
pushed away from it. If you hang suspended on a rope you are
pulling the rope downwards, but the rope is also pulling you
upwards. As you remain in the same position the two forces
must be equal and act in opposite directions.

We talk about a magnet attracting a piece of iron, because
usually the magnet is fixed and the iron free to move, but if we
arranged that both were free to move, the magnet and iron
would move towards each other. If the iron were fixed and the
magnet free to move, the magnet would travel towards the
iron. When you jump, you push down on the ground. When
a shell is fired from a gun the latter jumps back a little. At
the moment of firing there is a thrust between the shell and the
gun. The shell goes forward and the gun is propelled in the
opposite direction with equal force.

INERTIA

A body at rest will not move unless some force acts on it. If
we see a body at rest begin to move, we assume that some
force has acted on it. A body, moving in a straight line without
altering its speed, will continue to do so unless acted on by
some force. If we see a body moving in a straight line at

constant speed begin to increase or decrease its speed, or to change its direction, we assume that some force has acted on the body. The property that all bodies have of remaining in the same state of rest or motion unless acted on by some force is called the *inertia* of the body.

Our ordinary experience supplies us with many examples of this property of bodies. When you are in a car, your body is travelling at the same speed as the car. If the car suddenly stops, your body tends to go on. If you are standing up when this occurs, you fall in the direction in which the car was moving because your feet are prevented from going forward while the rest of your body is free to move. For the same reason, if you are standing in a train facing the engine, you tend to fall backwards if the train suddenly starts.

Suppose that a train is travelling at 60 miles an hour and something is thrown out of the window. As the object leaves the train it has a speed of 60 miles an hour in the same direction as the train is travelling. When it gets out of the window it loses a little of this speed owing to the resistance of the air, and so strikes the ground behind the window from which it was thrown. Nevertheless it hits the ground travelling forwards at a speed not much less than 60 miles an hour.

Suppose while travelling in a train you jump upwards from the floor of the carriage. You would expect to come down on the floor on the spot from which you jumped, and so you would. During the time you were in the air, out of actual contact with the train, the latter has travelled a considerable distance, and so have you. Although you were out of contact with the train, you had the same speed of the train owing to your inertia.

Expt. 215. Place a penny on a postcard resting on the rim of a tumbler. Sharply strike the edge of the card horizontally. The card leaves the tumbler, but the penny falls in. The penny, owing to its inertia, remains in its original position until deprived of the support of the card, when it falls into the tumbler.

Consider a motor bicycle attached to a sidecar travelling quickly in a straight line. Let the driver now turn sharply to the left. Owing to its inertia the sidecar tends to continue in a

straight line. This it is unable to do because it is fastened to the cycle. Under these conditions there is a tendency for the side-car to lift. This is well known, and drivers allow for this when turning corners. In sidecar racing the occupant of the side-car leans right over to prevent this lift. There is the same tendency to lift in a car, but it is not so pronounced. People who talk about 'going round a corner on two wheels' are usually exaggerating. If you ask them which two wheels, they usually say the wrong two unless they have studied mech-anics.

WEIGHT

Hold a large stone in your hand with your arm stretched horizontally. You will be con-scious that your muscles are exerting a force to keep the stone in position. Now with-draw your hand from the stone. The stone falls to the ground. From a state of rest it moves to the ground. Now we have seen that bodies do not move from a state of rest unless acted on by some force. Therefore the stone must have been acted on by a force. The force was its weight. *Weight is a force.*
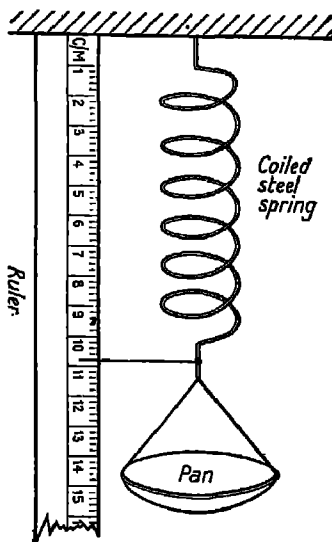
Fig. 128.

Expt. 216. Obtain a coiled steel spring. Stretch it a little by pulling the two ends apart. Note the muscular effort involved. You have to exert a force to stretch the coil.

Suspend the coil as shown in Fig. 128 and arrange a ruler vertically by its side. To the lower end of the coil attach a pan or tin box which should be heavy enough to stretch the coil a little. Note on the ruler the position of the bottom of the coil. For convenience an indicator as shown in the figure may be attached.

Put different weights in the pan and note the extensions.

Tabulate your results as follows:

No.	Weight in pan	Scale reading	Extension of the coil

When you have a 5 gm. weight in the pan, the coil is extended a certain amount. This extension is caused by a force of 5 gm. weight. The extension caused by a force of 10 gm. weight is double that caused by the 5 gm. weight. A spring-balance (Fig. 129) may be used for weighing objects and for measuring forces. Examine a spring-balance and test its accuracy by weighing some 'weights'. If you can obtain a suitable spring, try to make a spring-balance for yourself. Spring-balances are very convenient to use, but most of them are not very sensitive and are unsuitable for very accurate work.
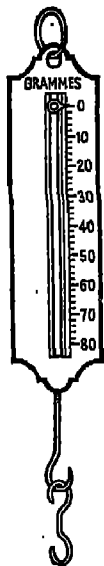
Let us consider what causes a stone to fall to the ground when the hand is removed. We do not know what makes it fall, but we do know that the stone behaves like all other stones in the same circumstances. In some way or other it is attracted to the ground. There must be a force which causes the stone to move towards the ground when free to do so. This force is called the *force of gravity* or *gravitational attraction*. (Note again that naming the force we know exists is not explaining it.) ISAAC NEWTON (1642–1727) showed that this force of gravity operated between all bodies in the universe. He did not explain why this force acted, but he stated the law according to which it acted.

Fig. 129.
Spring-balance.

MASS AND WEIGHT

The amount of matter in a body is called its *mass*. There is twice as much matter in two pounds of water as there is in one pound of water. One pound of lead contains the same amount of matter as a pound of water. The mass of a body is proportional to its weight. *Mass is not the same as weight.* The *weight of a body* is the *force* by which it is attracted to the earth. The *mass of a body* is the amount of matter in it. Thus the mass of a body will always be the same, whatever the position of the body. The weight, as we shall see, may vary. The spring-balance measures forces, not masses, but if two masses have equal weights, as shown by the spring balance, in the same place, then the two masses are equal.

FRICTION

Slide a smooth large stone along the smoothest surface you can find, e.g. ice. The stone will probably travel a long way before coming to rest, but, however far it travels, it will eventually stop. Something must have stopped it. We know that some force must have acted on the moving stone in order to bring it to rest. Repeat the experiment on a surface not quite so smooth, e.g. a paved foot-path. The stone stops after travelling a shorter distance. Hence the force bringing it to rest must in this case be greater. The difference is that in the first experiment the surface was smoother than in the second. The force that helped to stop the moving stone was the friction between the two surfaces. The friction was greater in the second case. *Friction is a force.* There was also friction between the moving stone and the air, but this would be about the same in both cases at the same velocity.

Friction between surfaces is often inconvenient, and makers of machines try to eliminate it as far as possible. It cannot be entirely eliminated, but the use of oil between surfaces helps. The use of ball-bearings also reduces friction. We shall return to this important subject when we come to the study of machines and work (p. 360).

On the other hand friction is essential. You have often

noticed how difficult it is to walk on very smooth ice. That is because there is very little friction between the ice and the shoes. If there was no friction between the surface of the road and our feet, we could not walk at all. A locomotive with a heavy load sometimes has difficulty in starting because the rails are smooth and the wheels 'skid'. This is remedied by putting sand on the rails to increase the friction between the two surfaces.

Friction is employed to stop a bicycle by the application of brakes. A bicycle running along a level road would eventually stop because of the friction between the moving parts of the bicycle and between the bicycle and rider and the air. If it is required to stop the bicycle before these forces do so, it is necessary to apply the brakes. This introduces friction between the brake-blocks and the rim of the wheel, which brings the bicycle to rest. It is necessary to apply a force to the pedals if it is required to keep a bicycle moving along a level road. The force applied overcomes the forces of friction. We shall study later what happens to a bicycle as it moves up or down hill.

THE FORCE OF GRAVITY

Newton showed that there is always a force of attraction between two masses. This force depends on: (1) the distance between the masses, and (2) the magnitude of the masses. In the case of a stone near the surface of the earth, the earth attracts the stone and the stone attracts the earth. They therefore move towards each other. The mass of the earth, however, is very much greater than that of the stone, and so the stone moves a correspondingly greater distance than the earth. For all practical purposes the earth's movement from causes such as this is so small that we need not take account of it.

Newton also showed that the attraction of the earth for bodies outside it, was the same as if the whole mass were concentrated at the centre. The earth is roughly a sphere, and everything on the surface of the earth or within the earth's influence, is attracted towards the centre of the earth. When some children first learn that the earth is like a ball, they wonder why the people on the other side of the earth do not

fall off. They cannot fall off because they are attracted to the centre of the earth. They are 'upside-down' compared with us, but they are not 'upside-down' compared with the centre of the earth. There is no 'up' and no 'down' except with respect to the centre of the earth. 'Down' means towards the centre of the earth, and 'up' means the opposite direction.

The nearer two bodies are to each other the greater is the attraction due to gravity. The weight of a man in an aeroplane which is still on the ground is greater than the weight of the same man in the aeroplane when it is (say) two miles high. This is because the man's weight is the force due to the earth's attraction, and the force depends on the distance of the man from the centre of the earth. The farther the man is from the centre of the earth the less is the force of attraction. Of course the man's mass remains the same wherever he is.

The earth is not a true sphere. The diameter measured through the poles is slightly less than that measured through the plane of the equator. This means that as a man travels from the equator to one of the poles, he is gradually getting nearer to the centre of the earth. The force of attraction on his body will therefore increase as he travels towards a pole. A man's weight at the poles is therefore greater than at the equator. This would be shown if he were weighed on a spring-balance, if a sufficiently sensitive one existed, because this measures forces. It would not be shown if he weighed himself on an ordinary balance like a pair of scales, because the 'weights' are affected in the same proportion as the weight of the man by the variation in the force of gravity.

We have considered gravitation from the point of view of how it affects bodies on or near the surface of the earth. It must not be forgotten, however, that every particle of the universe attracts every other particle. Some are so far away from each other that for all practical purposes there is no result. Newton discovered the law of gravitation when studying the movements of the heavenly bodies. If all bodies attract each other, why are they not by this time as near to each other as possible? Why are they not all in contact? The reason is that there are other factors operating besides that of gravitation.

HOW CAN FORCES BE MEASURED?

We have seen on page 304 that a force is that which tends to change the state of rest or motion of a body. How are we to measure forces? What do we mean when we say that one force is greater than another? Suppose we apply a force to a football by kicking it. If the ball is already in motion, the force applied makes it travel more quickly or more slowly or changes its direction. The bigger the kick (that is, the greater the force applied) the greater is the effect on the motion of the ball. We can compare the forces applied (the strengths of the kicks) by comparing the effects of the forces on the ball. We assume that if a force has twice the effect of a second force, the first force is double the second. By comparing the effects of forces on moving bodies, we can compare the forces which cause the effects. Before we can choose a unit of force with which to compare other forces, we must learn something more about moving bodies.

VELOCITY

We can describe the velocity of a body by stating how far it travels in a certain time. We say that a motor-car is travelling at 30 miles per hour, or 44 ft. per second. A car travelling at 60 miles an hour has a velocity double that of one travelling at 30 miles an hour. Velocity is measured by the distance travelled in unit time. The *unit of velocity* is the velocity of 1 ft. per second, in the English System, and 1 cm. per second in the Metric System.

If the velocity of a body remains the same over a certain distance, it is said to move with *uniform velocity* over that distance. Nearly all moving bodies, however, do not move uniformly. The velocities change from time to time according to the forces acting on the bodies. When we say that a motor-car is travelling at 30 miles an hour we do not mean that it has travelled for one hour with uniform velocity and has covered 30 miles. A car may travel one mile (or one yard) at the rate of 30 miles an hour. We know from our experience that cars do not travel far with uniform velocity. In the course of a journey a car
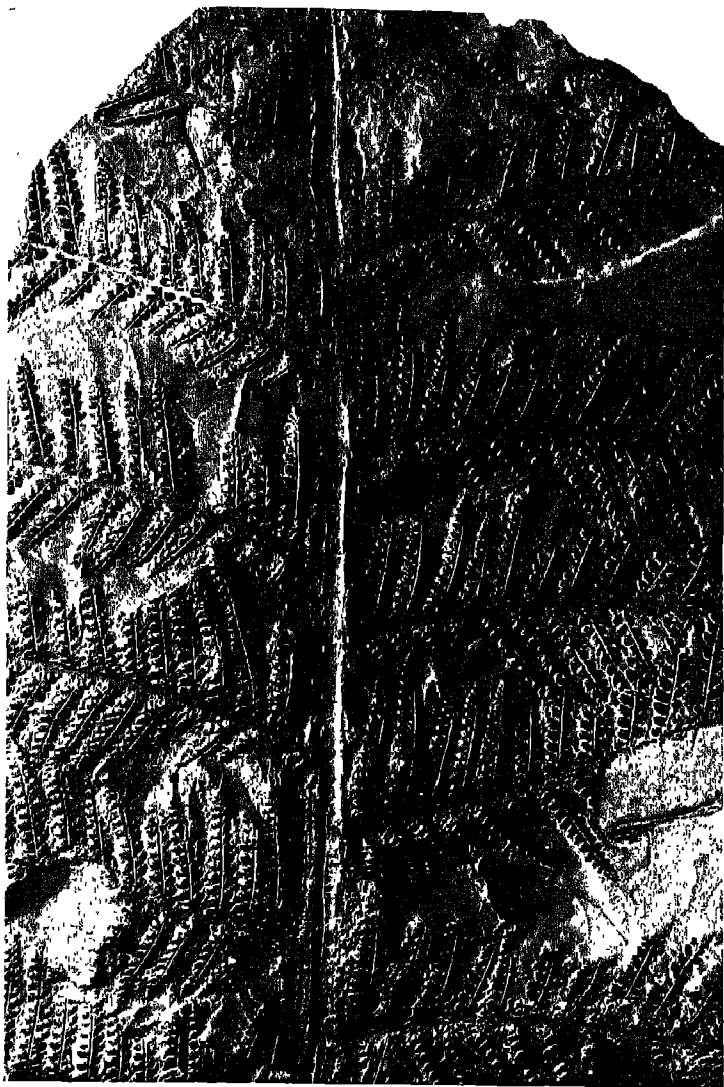
V. A fossil plant from the coal measures

VI. Freshly drawn salt heaped on a draining platform ('hurdle')
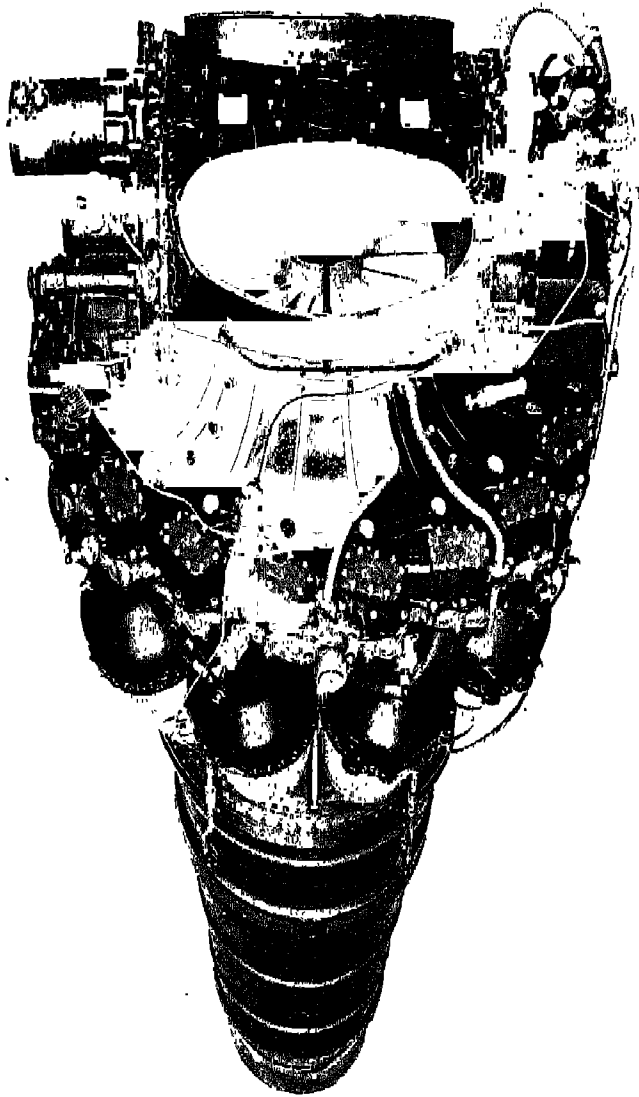
VII. The De Havilland Ghost gas turbine engine

VIII. A trolley bus being tested for stability

might travel at the rate of 60 miles an hour for part of the time and then it might travel a further distance at 10 miles an hour. Its velocity would vary. If, however, it completed 30 miles in one hour we would say that the car did the journey at the rate of 30 miles an hour. This is really the *average velocity* of the car for that journey.

What exactly do we mean when we say that a car passed us at 30 miles an hour? We mean that at the instant it passed us it was travelling at such a velocity that if it had continued to travel for an hour at that velocity it would have travelled 30 miles. The velocity at any particular instant may be taken as the average velocity during a very small interval of time in the neighbourhood of that instant.

Distance travelled = Average velocity × Time.

ACCELERATION

When a moving body increases its velocity it is said to accelerate. Consider a motor-car starting from rest. Suppose that at the end of the first second it has reached a velocity of 5 m.p.h., at the end of the next second a velocity of 10 m.p.h., and at the end of the third second a velocity of 15 m.p.h. Each second the velocity increases by 5 m.p.h. That is to say, the rate of increase of velocity is 5 m.p.h. per second. It is important to realize that an acceleration cannot be stated without mentioning time twice. Acceleration is not the change in velocity, but the *rate* of change of velocity. If a train changes its velocity from 50 m.p.h. to 60 m.p.h. in 1 minute, the change in velocity is 10 m.p.h. but the acceleration, that is the rate of change of velocity, is 10 m.p.h. per minute. It is usually convenient to use the same unit of time in both cases, and instead of saying an acceleration of 10 m.p.h. per minute we can say an acceleration of 0·24 ft. per sec. per sec.

The *unit of acceleration* is 1 ft. per sec. per sec. in the English System, and 1 cm. per sec. per sec. in the Metric System.

The velocity of a body may decrease. The rate of decrease of velocity is called the *retardation*. This may be looked upon as a negative acceleration. When the velocity of a body increases or decreases by equal amounts during equal intervals of time the

As a result of his experiments Galileo discovered that *the distance travelled down the plane varies as the square of the time.*

Let us repeat Galileo's experiment.

Expt. 220. Arrange an inclined plane from 8 to 10 ft. long with a shallow groove along which a metal ball can roll (Fig. 130). A large ball-bearing is suitable for the purpose. Graduate the board clearly in feet. Place a piece of wood as a stop at the 1 ft. graduation. Hold the ball at the 0 graduation and have a stop-watch in the other hand. Release the ball and start the stop-watch at the same time. When the ball hits the wooden stop press the stop-watch. Note the time. Repeat
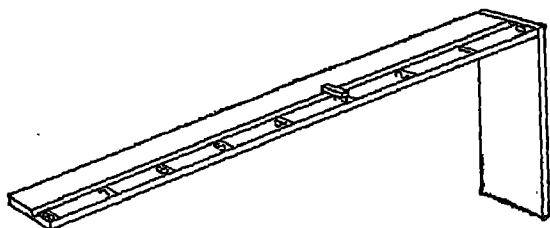
Fig. 130. Galileo's experiment.

the experiment several times at this graduation until you obtain three consecutive readings which do not differ much. Take the average of these three readings for the time. Repeat the experiment at every foot graduation taking the average as before. Tabulate your readings as below:

Distance in feet (s)	Time in secs. (t) (Average of last three consecutive observations)	t² (Calculated or from tables)
1		
2		
&c.		

Draw a graph plotting *s* against t^2. You will find that the graph is a straight line through the origin (allowing for experimental error) which means that $s \propto t^2$.

The same result is obtained if we observe the movement of two bodies connected by a string passing over a pulley. One body falls as the other rises but by choosing the weights of the bodies we can arrange for the falling body to fall at such a rate

that its time of fall can be observed. If we allowed a body to fall freely it would fall so quickly that it would be impossible to time its fall accurately. Atwood's machine is a convenient apparatus for performing this experiment but any similar arrangement which will enable the time of fall to be measured accurately may be used.

Expt. 221. Atwood's machine (Fig. 131) consists of a pulley, *A*, fixed at a height above the ground. The wheel must turn as freely as possible so some arrangement must be made to reduce friction at the axle. Two equal weights, *B* and *C*, are attached to a string which passes in the groove over the pulley. Weight *C* rests on a hinged platform, *D*, which can be removed at will allowing the weight *C* to fall. *C* is surmounted by another weight, *E*, in the form of a rider. *F* is a movable ring through which the weight *C* is able to pass but which stops the rider *E* which is wider than *C*. *G* is a platform on to which *C* can fall. The support of *A* is graduated in centimetres (or inches) so that the vertical distance from *C* to *F* can be accurately known.

Fix the ring *F* in position and place the rider on the weight *C*. Note the distance between *D* and *F*. As you press the stop-watch let the hinged platform *D* fall away thus allowing weights *C* and *E* to fall. Take the time as the rider *E* hits *F*. Repeat the experiment until three consecutive observations differ by very little and take the average of these for the time of fall. Repeat the experiment for different positions of *F*. Tabulate your results and draw a graph as you did in the previous experiment. You will obtain the same result, $s \propto t^2$.
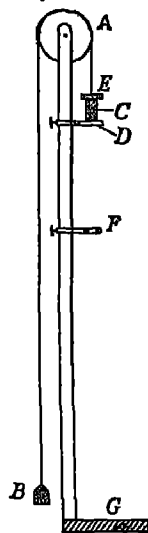
FIG. 131. Atwood's machine.

This is what Galileo found, but he used the inclined plane method. For example: Suppose a ball rolls 1 ft. down the plane in the first second. In 2 seconds it rolls 4 times (2^2) as far. That is, in 2 seconds it rolls 4 ft. (1 ft. $\times 2^2$). In 3 seconds it rolls 3^2 times the distance it rolls in the first second, that is 9 ft. (1 ft. $\times 3^2$), and so on. What is true for the inclined plane is also true for a body falling freely. The distance fallen varies as the square of the time.

It has been found that when a body falls freely from a height of 16 ft. it reaches the ground in 1 second. If it falls

The Impact of Two Inelastic Bodies Moving in the Same Line

We need an apparatus which will enable us to study the impact of different masses moving with different velocities in the same line. This may be done with a ballistic balance.

Expt. 222. A and B are two long pendulums (Fig. 132) supporting two platforms C and D from a framework. C and D may be pulled aside, by means of the strings S_1 and S_2 passing through the rings R_1 and R_2, and their distances from the centre can be measured on the horizontal scale E. These measurements are facilitated by pointers fixed to the inside edges of C and D. Projecting pins and a layer of cork are fitted to the platforms so that when they collide, on releasing the strings, the pins stick into the cork and the platforms remain together acting as inelastic bodies. The masses of C and D must be known so that when various masses are placed on the platforms the total mass of each may be known. We need to know also the velocities of C and D at the instant of impact. It can be shown that, *when the pendulums are long compared with the horizontal distances moved the velocity at the lowest point is proportional to the distance travelled.* That is, if C moves 20 cm. and D 10 cm. then the velocity of C at the moment of impact will be twice that of D. We can, then, vary the velocities of the colliding bodies by displacing the platforms different distances. That is Velocity $= A \times$ the distance displaced, where A is a constant. The Momentum $=$ Mass $\times$ Distance displaced $\times A$. It is never necessary to find A and for convenience some people leave it out altogether, but do not attempt this until you are sure you understand why it is possible to do so.

Fig. 132. Ballistic balance.

1. Make the total mass of C 100 gm. and D equal to C. Pull C and D back equal distances (20 cm.). Release the string. C and D collide and remain together and stationary.

Momentum $=$ Mass $\times$ Velocity (p. 319).
Momentum of $C = 100 \times 20A = 2{,}000A$ in one direction.
Momentum of $D = 2{,}000A$ in the opposite direction.
Total Momentum of C and $D = 2{,}000A - 2{,}000A$
$= 0$ before impact.

As the combined bodies C and D remain stationary after impact the momentum is $200 \times 0 = 0$. That is, the momentum before and after impact is the same, in this case 0.

2. Make C 100 gm. and allow it to remain at 0 on the scale. Let D be 200 gm. and move from 20 cm. to the right. Note that after the impact C and D move together to the left for about 13 cm.

Momentum before impact, $C = 0$ and $D = 4,000A$ (i.e. $200 \times 20A$). Total momentum towards the left $= 4,000A$.

Momentum after impact $= 13A \times 300 = 3,900A$, which, allowing for experimental error, is equal to $4,000A$, the momentum before impact.

3. Repeat these experiments with different masses for C and D and from different distances from the centre (i.e. with different velocities at the instant of impact), and note the results, comparing the total momentum before impact in one direction and the momentum of the combined bodies after impact.

Your results will show, within the limits of experimental error, that: *when two inelastic bodies moving in the same line collide, the total momenta before impact in one direction is equal to the momentum of the combined bodies in the same direction after impact.* This fact is known as the Law of Conservation of Momentum.

Let a mass, M_1 gm. move to the right with a velocity U_1 cm. per sec. and let it collide with a mass of M_2 gm. moving to the left with a velocity of U_2 cm. per sec. and suppose the combined masses $(M_1 + M_2)$ to move together to the right with a velocity of V cm. per sec.

Before impact,

Momentum of $M_1 = M_1 U_1$ measured to the right $(+)$
 „ „ $M_2 = M_2 U_2$ „ „ „ left $(-)$
Total momentum of M_1 and $M_2 =$
$$M_1 U_1 - M_2 U_2 \text{ to the right.}$$

After impact,

Momentum of combined bodies to the right $= (M_1 + M_2) V.$

Then, according to the Law of Conservation of Momentum,

$$M_1 U_1 - M_2 U_2 = (M_1 + M_2) V.$$

Example 1. It is possible to find the weight of a body by the use of a ballistic balance but it is far from being an accurate

method. It is interesting theoretically. Suppose C (in Fig. 132) is 10 lb., and D of unknown weight. Allow C and D to collide starting 12 inches from the middle of the scale, and suppose C and D continue together to the right. Then,

Momentum of $C = 10 \times 12A = 120A$ to the right.

„   „ D $= 12DA$ to the left.

Total momentum $= (120 - 12D)A$ to the right.

Suppose that after impact the combined masses C and D travel 5 inches to the right.

Then   Momentum $= (10 + D) \times 5A$ to the right.

Then $(120 - 12D)A = (10 + D) \times 5A$.

Whence $D = 70/17 = 4$ lb. (approx.).

Example 2. A railway truck weighing 5 tons and travelling at 40 feet per sec. collides, without rebounding, with a stationary truck of 6 tons. At what speed do the trucks continue in the same direction?

Let V be the velocity after impact.

Momentum before impact $= (5 \times 40) + (6 \times 0) = 200$.

Momentum after impact $= (6 + 5)V$

Then, $11V = 200$, i.e. $V = 18 \cdot 2$ ft. per sec. (approx.).

Example 3. Two wooden balls, 5 oz. and 7 oz., collide while moving from opposite directions with velocities of 10 ft. per sec. and 20 ft. per sec. respectively, and do not rebound. What is the velocity after impact?

Let V be the velocity after impact.

Momentum before impact in one direction
$$= (7 \times 20 - 5 \times 10) = 90.$$
Momentum after impact in the same direction
$$= (5 + 7)V = 12V.$$

Then $12V = 90$.

.i.e. $V = 7\frac{1}{2}$ ft. per sec. in the direction followed by the 7 oz. ball.

Example 4. A converse problem is the recoil of a gun when a shell is fired. Before the shell is fired the shell and gun are at rest together. On firing, the momentum of the shell must be equal to that of the gun. If a gun weighs 20,000 lb. and a shell, weighing 500 lb., leaves the gun with a velocity of 1,500 ft. per sec., then,

Momentum of shell $= 500 \times 1,500 = 750,000$.

,,   ,, gun $= 20,000 \times V$ where V is the velocity at which the gun starts to recoil.

Then, $20,000V = 750,000$

$V = 750,000/20,000 = 37\frac{1}{2}$ ft. per sec.

MEASUREMENT OF FORCE

On page 312 we saw that we could compare forces by noting the effects of forces on moving bodies. When a force alters the motion of a body it does so by speeding it up or by slowing it down. A force acting on a body free to move causes an acceleration. We can compare forces by noting the acceleration produced by different forces on the same body.

Acceleration ∞ Force.

That is,   Rate of change of velocity ∞ Force.

Now we have seen that a change of velocity, the mass being constant, must mean a change in momentum.

Therefore, Rate of change of momentum ∞ Force.

Or,   Mass $\times$ Rate of change of velocity ∞ Force.

That is,   Mass $\times$ Acceleration ∞ Force.

Units can be chosen so that Mass $\times$ Acceleration $=$ Force

or $M.a = F.$

The *unit of force* is that force which will produce an acceleration of 1 cm. per sec. per sec. in a mass of 1 gm. This is the C.G.S. unit and is called the *dyne.* The unit in the British System is called the *poundal.* It is the force which will produce an acceleration of 1 ft. per sec. per sec. in a mass of 1 lb.

The dyne and the poundal are called *absolute units.* Other

units are in common use, especially for forces acting on bodies when no motion is caused. They are called *gravitational units*. In the C.G.S. system the unit is the 1 gm. weight. It is the force due to gravity on a mass of 1 gm. It produces an acceleration of 981 cm. per sec. per sec. As a dyne produces an acceleration of 1 cm. per sec. per sec. in a mass of 1 gm., clearly 1 gm. weight is equal to 981 dynes.

In the British system the gravitational unit of force is the 1 lb. weight. As this force produces an acceleration of 32 ft. per sec. per sec. in a mass of 1 lb., then 1 lb. weight is equal to 32 poundals.

THE SIMPLE PENDULUM

When Galileo carried out his experiments on falling bodies he had to devise some means of measuring small intervals of time, because in his day there were no clocks and watches such as we have to-day. He did this by allowing water to drip through the hole in the bottom of a large tank of water. By weighing the water which dripped he was able to compare intervals of time. Before he died Galileo solved the principle of the pendulum and thus made possible the construction of more accurate mechanical clocks. When in the Cathedral of Pisa one day in 1581, Galileo's attention was attracted to a lamp-holder swinging from the roof. He timed the swings to and fro by counting the beats of his pulse. He noticed that although the swings became narrower, the wide and narrow swings took the same time. This led him to study the pendulum and the designing of a pendulum clock.

Expt. 223. Fit up a pendulum as shown in Fig. 133. It consists of a long thread supporting a metal ball. The method of suspension is important. Pass the thread through the halves of a split cork. Clamp these tightly together and support on a retort stand. This is to ensure that the pendulum is suspended from one point and that the point of suspension does not change as the pendulum swings. Mark the position of the bob when at rest. Allow the pendulum to swing. The *amplitude* of the swing is the distance from the central position to either extreme. A complete swing is the motion of the pendulum bob from the central position until the next passage of this position in the same direction. This is called one *oscillation*. (See vibration, p. 470.)

What is the effect of varying the amplitude? Hold the bob at rest a very little out of the vertical. Release the bob, allowing it to swing. Press the starter of a stop-watch when the bob passes the mark. Count the swings. At the end of (say) the fiftieth complete swing stop the watch. This will give the time of 50 complete swings. Find by division the time of one complete swing. Repeat the experiment several times increasing the amplitude each time, but never allowing it to become large. The time of swing, sometimes called the *period*, is the same for all amplitudes with the same pendulum. A motion which repeats itself is called a *vibration*. The movement of the bob of a long pendulum is a very simple kind of vibration (p. 470).

What is the effect of varying the weight of the bob? Use the pendulum in the same way with small amplitudes, but substitut-ing bobs of different weights.
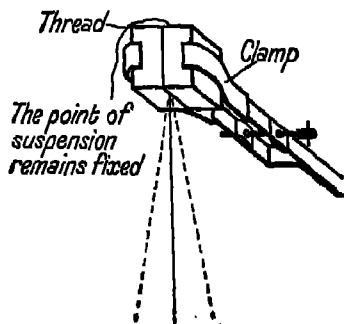
Fig. 133. How to suspend a pendulum.

Find the times of swing for bobs of different weights. You will find that the weight of the bob has no effect on the time of swing.

What is the effect of a change in length of the pendulum? The length of the pendulum is from the point of support to the centre of the bob. Find the diameter of the bob by using a pair of callipers. Measure the length of the thread from the point of support to the bob and add to this length the radius of the bob. Allow the pendulum to swing as in the previous experiments and from the time of 50 swings calculate the time of one swing. Repeat for different lengths of the pendulum. The latter can be conveniently shortened by pulling the thread through the clamping arrangement. You will find that the shorter the pendulum, the shorter is the time of swing. Tabulate your results thus:

Length of pendulum in cm. (L)	No. of swings	Time in secs.	Time of one swing (t)	$\sqrt{L}$	$\dfrac{t}{\sqrt{L}}$

You will find that, within the limits of experimental error, the ratio $t/\sqrt{L}$ is constant. Let this constant be represented by k.

Then, $$\frac{t}{\sqrt{L}} = k.$$

That is, $$t = k\sqrt{L},$$

or, the time of swing varies as the square root of the length of the pendulum.

It can be proved that the time of swing is given by the formula

$$t = 2\pi \sqrt{\frac{L}{g}},$$

where L is the length of the pendulum and g the acceleration due to gravity.

THE SECONDS PENDULUM

Each oscillation of a pendulum consists of two *beats*, one each way. A seconds pendulum is one in which the time of one beat is one second. The time of the complete swing will therefore be two seconds. Using this value in the formula we can calculate the length of the seconds pendulum.

$$2 = 2\pi \sqrt{\left(\frac{L}{32 \times 12}\right)}$$

where L = length of the pendulum in inches.

That is, $$\frac{7}{22} = \sqrt{\left(\frac{L}{32 \times 12}\right)}.$$

Therefore, $$\frac{49}{484} = \frac{L}{32 \times 12}.$$

Therefore, $$L = \frac{49 \times 32 \times 12}{484}$$

$$= 38 \cdot 9 \text{ inches.}$$

A pendulum about 39 inches long is used for controlling the speed of a Grandfather Clock (Fig. 134). Examine a pendulum clock and try to find out how the pendulum controls it. There is a toothed wheel which, if free, would turn very quickly. There is an arrangement by which the pendulum allows one tooth of the wheel to escape at each beat of the pendulum, that is, one every second.

If a pendulum clock is not keeping good time it may be regulated by altering the length of the pendulum. A nut travelling on a screw thread is provided for this purpose under the bob. If the clock is losing time the nut is turned to the right thus raising the bob a little and so making the pendulum shorter. This reduces the time of swing. If the clock is gaining time, the bob is lowered slightly.
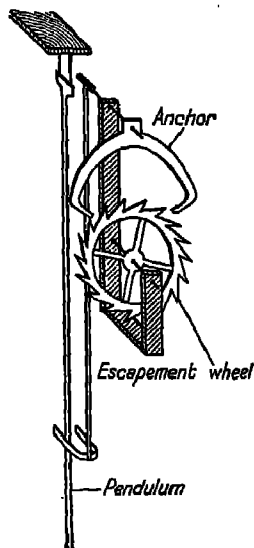
Fig. 134. The pendulum of a grandfather clock.

QUESTIONS

1. What is the function of the pendulum of a clock and how is this function performed?

2. Find the time of swing of a pendulum 8 ft. long.

3. Find the length of a pendulum which makes 50 complete swings in 70 seconds.

4. A train does a journey of 200 miles with an average velocity of 45 m.p.h. It does the first 100 miles in 2 hours. Find the average velocity for the next 100 miles.

5. A cyclist reached the top of a downward slope when he was travelling at 10 m.p.h. The slope gave him an acceleration of 2 ft. per sec. per sec. In what time would his speed have doubled?

6. A body 10 lb. weight falls freely for 100 ft. With what momentum does it hit the ground?

7. What do you understand by the term *force*? Illustrate your answer by reference to weight and friction. How is force measured?

8. Why is it a bad practice for passengers to stand in a carriage when a train is about to stop at a station? If you had to jump off a moving bus what precautions would you take? Give your reasons.

9. Define the term acceleration. What do you understand by the term 'acceleration due to the force of gravity'?

10. When a body moves (*a*) from the equator towards one of the poles or (*b*) from the surface of the earth to a point above the earth, the weight changes but the mass remains the same. Explain this.

XX

MACHINES

As animals grow, they quickly learn to use their limbs for getting about and for other purposes. A man's hands are wonderful instruments by which he can perform many operations of great complexity. Man, however, is not, by his unaided efforts, able to move objects of any great weight. If he had to rely on his own efforts, he could not build the houses he does, and most of the engineering feats of the present day would be impossible. You have no doubt watched men at their work and noted the devices they employ to help them. When we think of the machines which man has invented, we should think not only of complicated machines like printing machines, steam engines, and motor-cars, but also of less complicated machines like lawn mowers, bicycles, pulleys, fire tongs, spades, barrows, crowbars used as levers, and even the stairs by which he ascends to his bedroom.

Many of the machines man uses enable him to exert a large force by employing a smaller one, but this is not always the case. The applied force is usually called the *effort*, and the resisting force the *load* or *resistance*. In some machines the effort is more than the load. We shall see that these machines have other advantages.

A machine is any contrivance by which an effort applied at one point and in one direction is made to overcome a resistance acting at another point and in another direction. We shall study some simple machines. Complicated machines are combinations of simple machines.

THE LEVER

Expt. 224. Obtain a crowbar or a rod of wood which will not easily bend. Place one end of it under a heavy object such as a large stone or a heavy piece of furniture. Place a support for the crowbar a short distance from the object to be raised (Fig. 135). Press down on the free end of the bar. The heavy object rises. Try to raise the object without the aid of the bar. Use the bar again, but this time

press down, not on the free end of the bar, but some distance from the end. Was it more, or less, difficult? Repeat, pressing down still farther from the end.

You find that the easiest way to raise the weight is to press down at the end of the bar. If you tried the experiment again with a longer bar you would find it still easier. This machine is called a lever. Let us see how it works.
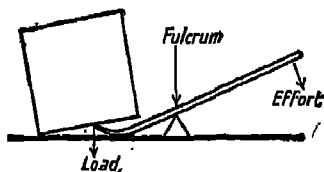
Fig. 135. A crowbar used to raise a heavy load.

The point of support about which the lever turns is called the *fulcrum*. When you play on a see-saw you are using a lever to raise your friend at one end of the plank. Consider a plank balanced and turning freely on the trunk of a fallen tree. One boy sits at one end. Where shall the other boy, who we will suppose is the heavier, sit? You know from your play that if the heavier boy sits at the end he will go down and will not rise unless he pushes the ground vigorously with his feet. If the heavier boy, however, sits nearer to the tree trunk he can find a position where his weight will just balance the lighter weight of the other boy. The see-saw can then be enjoyed without too much help from the feet. In your play you have probably found another way of arranging the see-saw so that both boys can sit at the end of the plank. In this case you have to alter the position of the tree trunk so that the distance of the lighter boy from the trunk is greater than that of the heavier boy. There is obviously some relation between the weight and the distance on the one side and the same quantities on the other. Let us try to find experimentally what this relation is. You could do this on a see-saw if you weighed a number of boys and marked your plank into feet, but we can get a more accurate result by con- structing an artificial see-saw.

MOMENTS

Expt. 225. Fit up the apparatus as in Fig. 136. *AB* is a long ruler. It is balanced on a thin peg passing through a hole in the middle of the ruler. The peg acts as a fulcrum.

If the ruler does not balance exactly when the mid-point is at the fulcrum, make it do so by increasing the weight of one side by dropping a little paraffin wax or sealing wax on the end or by placing a folder clip on one side. For convenience use a ruler marked in centimetres and millimetres (but any units will do). Tie pieces of thread to various weights so that they may be suspended from any part of the ruler by a loop of the thread. Suspend two unequal weights one on each side of the fulcrum. Move them about until the ruler is balanced. Note the distance of each from the fulcrum. Using the same two weights find other positions in which the ruler balances. Repeat the experiment with other weights and tabulate your results as follows:
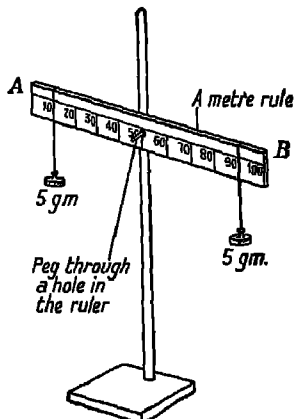
Fig. 136.

Left-hand side		Right-hand side		Product of the weight and distance from the fulcrum on each side	
Weight in grams (x)	Distance of weight from fulcrum in cm. (a)	Weight in grams (y)	Distance of weight from fulcrum in cm. (b)	$x \times a$	$y \times b$

You will find that the numbers in the last two columns are equal or very nearly equal.

Consider the ruler, when it is in a state of rest, having a weight suspended on each side of the fulcrum. The weight on the left is a force tending to turn the ruler about the fulcrum. The weight on the right is a force tending to turn the ruler in

the other direction about the fulcrum. As the ruler does not
turn at all it follows that the turning effects of the two forces
are equal. The *turning effect* (or *leverage*) on the left is the
product of the force and the distance from the fulcrum. On
the right it is the same. When the turning effects are the same
the ruler balances.

When a force acts so as to produce (or tend to produce) a
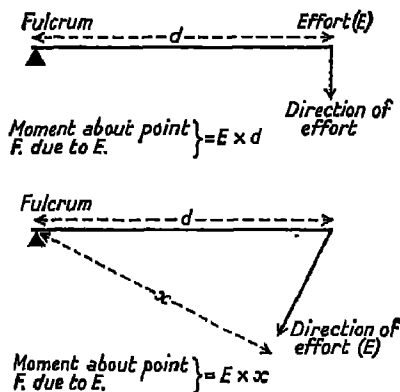
Fig. 137.

turning effect about a fixed point or axis, the turning effect is
measured by the product of the force and the *perpendicular*
distance of the point or axis from the line of action of the force
(Fig. 137). This turning effect is called the *moment of the force
about the point*. More briefly:

$$\text{Moment} = \text{Force} \times \text{Arm}.$$

THE USE OF THE LEVER

When an effort is used to raise a heavy stone by a crowbar,
the fulcrum is very near the stone and the effort is applied as
far as possible from the fulcrum. Consequently, the effort is
less than the load. In theory it is possible to raise any load by
a lever of this type. If the load is very great a very long lever
may be used. In practice, however, there is a limit to the load
which can be raised by a lever. The lever must not bend or

break. When the load is very great it is impossible to find a suitable rod that will not bend or break. The position of the fulcrum is also important. It may be impossible to find a suitable position for the fulcrum (which must not move). With a very heavy load the fulcrum may sink when the effort is applied and the load would not be raised.

A lever is useful for raising loads short distances. It is of no use for raising loads great distances.

MECHANICAL ADVANTAGE

It is important to know what advantage we are gaining when we use a machine. This is found by comparing the load raised with the effort applied.

$$\text{Mechanical advantage} = \frac{\text{Load}}{\text{Effort}}.$$

If the load is the same as the effort the mechanical advantage is 1. If the load is 10 times the effort the mechanical advantage is 10. The mechanical advantage may be less than 1. In that case the load is less than the effort. This is really a disadvantage mechanically, but there may be other advantages which make the method useful. We shall see (p. 338) that a very important lever in the human body has a mechanical advantage less than 1.

DIFFERENT KINDS OF LEVERS

So far we have dealt with only one kind of lever. We shall now deal with other types and give some common examples of each. Try to find other examples as you watch men at their work, and try to understand why they are used for that particular purpose. The study of the following examples will help you to do this.

There are three classes of levers:

I. Fulcrum between the load and the effort

This is the type we have already studied. The mechanical advantage may be greater than, equal to, or less than 1.

Examples.

1. The **crowbar** with the fulcrum on a block a little distance from the load (Fig. 135).
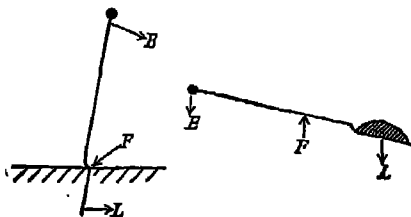
FIG. 138. A spade as a lever
(fulcrum between the load and the effort).
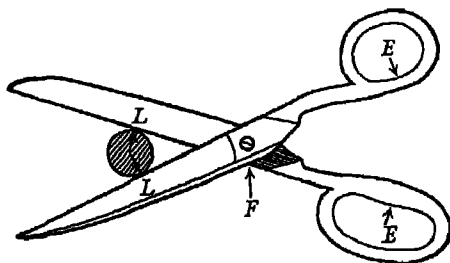
FIG. 139. Scissors, a double lever with the
fulcrum between the effort and the load.

2. A **poker** used to lift a fire, the fulcrum being a bar of the grate.

3. A **spade** used as shown in Fig. 138.

4. The **handle of a pump.**

5. The **see-saw.** The object of a see-saw is not to raise a load by means of a smaller effort, but to arrange that two (possibly different) weights should be balanced so that the plank will freely turn about the fulcrum.

6. A pair of **scissors.** Examine Fig. 139 and note that a pair of scissors is an example of a double lever of this class.

7. A **pick** (Fig. 140). The effort should be in the direction shown in the figure (i.e. at right angles to the line from the end of the handle to the ful-
crum) to obtain the greatest turning effect about the fulcrum, because the mo-ment is the product of the effort and the *perpendicular* distance from the fulcrum to the line of action of the effort.
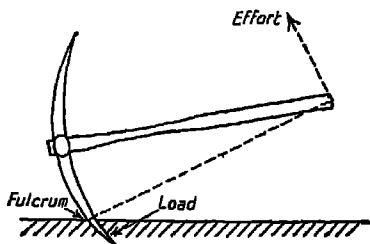
FIG. 140. A pick as a lever fulcrum between the effort and load.

8. The **balance** (Fig. 141). This is a very impor-tant application of the principle of the lever. If the arms are alike in all respects, including length, then there is a balance when equal weights
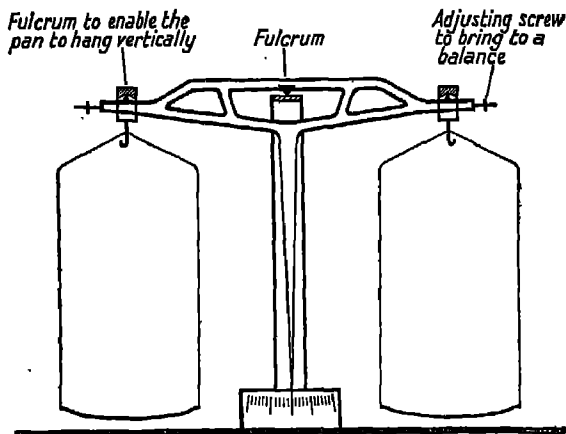
FIG. 141. A diagram of a balance to show the moving parts.

are placed on the pans. If an unknown weight is placed on one pan and balanced by known weights on the other pan, the unknown weight then becomes known. Consider very carefully what the balance does compare. Suppose there is a 1 gm. weight in the left-hand pan balanced by a piece of chalk
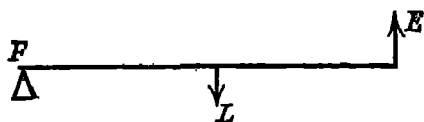
FIG. 142. Diagram of a lever
with the load between the fulcrum
and the effort.
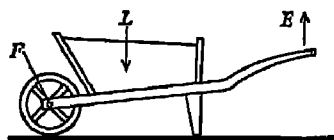
FIG. 143.
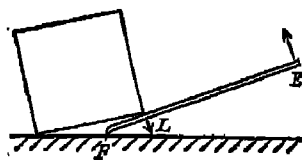
FIG. 144.
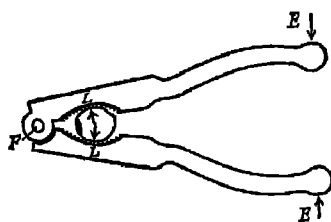
FIG. 145.
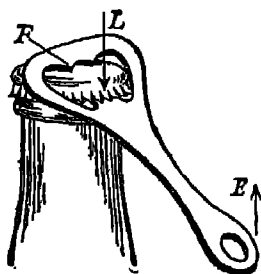
FIG. 146. A mineral bottle
opener.
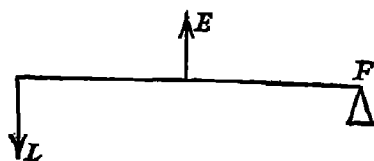
FIG. 147. Diagram of a lever
with the effort between the fulcrum
and the load.

in the right-hand pan. Both the piece of chalk and the 1 gm. weight are attracted to the earth by the same force due to gravity. This force varies in different parts of the earth, but the two bodies always balance no matter where they are. The balance really compares the mass of the piece of chalk with the mass of the 1 gm. weight. When measured in this way we can say that the mass of the piece of chalk is 1 gm., but its weight may be slightly different, and this would be shown on a spring balance if one could be made sensitive enough.

Expt. 226. Examine a balance such as is used in school. It is usually called a 'chemical balance'. Take it to pieces. (You had better do this under supervision as the parts are easily damaged.) Note the arrangement for the fulcrum and the device which enables the pan to hang vertically. Allow the balance to swing without any loads in the pans. When it comes to rest does the indicator show that both sides are equally balanced? If not, how can the balance be adjusted? Find out, by examination, how this is done by employing the principle of the lever. Note the two screws which in effect lengthen one arm compared with the other.

II. Load between the fulcrum and the effort (Fig. 142)

The mechanical advantage of this class is always greater than 1. This can easily be seen from the figure.

Examples:

1. The **wheelbarrow** (Fig. 143). The fulcrum is the axle of the wheel, the effort is at the end of the handles (half the effort applied to each handle), and the load is between the fulcrum and the effort.

2. A **crowbar** used as shown in Fig. 144. The fulcrum is on the ground, and the load between the effort and the fulcrum.

3. A pair of **nut-crackers.** This is a double lever of this class. Examine a pair. Find the position of the fulcrum, load, and effort (Fig. 145).

4. Opener for mineral water bottles closed by metal caps (Fig. 146).

III. Effort between the fulcrum and the load

Examine Fig. 147 and note that in this type the effort is greater than the load. The mechanical advantage is less than 1 and for this reason there are not many practical examples of

this type. A pair of tongs or tweezers is a double lever of this type (Fig. 148). One of its advantages is that it enables the effort to be applied at a more convenient place. Many of the levers in the human body and in the bodies of other animals are of this type. At first sight it appears strange that this should be so. Let us examine the use of levers in the animal's body by considering two examples from our own body.

THE MOVEMENTS OF THE FOREARM

Fig. 149 illustrates the bones of the human arm with two of the muscles which move the arm at the elbow. Raise your forearm and watch what happens to the large muscle on the front of the arm (the *biceps*). This muscle is attached by two tendons to the *scapula* (shoulder blade). At the other end it is attached by one tendon to the *radius* of the forearm a short distance from the elbow joint. When the muscle contracts and becomes thicker, the forearm is drawn up, turning about the elbow joint. It is not the raising of your arm that thickens the biceps muscle. On the contrary, the muscle thickens as it shortens and thus draws up the forearm. The load is the forearm, the hand, and anything it contains. The effort is applied between the load and the fulcrum, which is the elbow joint. This is an example of the third order. The effort required is greater than the load and so the mechanical advantage is less than 1. But a slight extension or contraction of the length of the muscle causes a large movement of the forearm and the hand. When the biceps shorten 1 inch the hand is raised about a foot. This arrangement is well adapted for the control of movement rather than for mechanical advantage. The biceps muscle is called a *flexor* muscle because it causes the flexion or bending of the forearm.

The straightening or extension of the forearm is brought about by an *extensor* muscle called the *triceps*. This is attached by a tendon to the scapula and by two other processes to the *humerus*. The other end is attached to the extreme end of the *ulna*. This extension of the ulna can easily be felt in the elbow. It is the 'funny bone'. About an inch below this is the fulcrum where the ulna is in contact with the humerus. Thus the effort
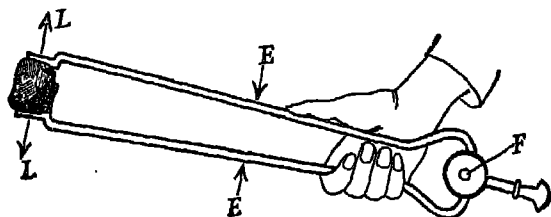
FIG. 148. Coal-tongs.
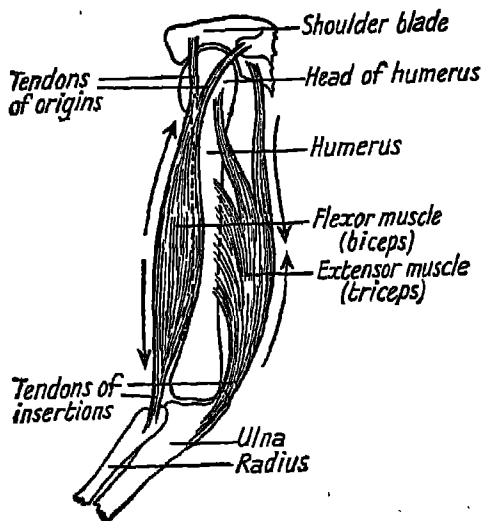
FIG. 149. The human arm.

and resistance are on opposite sides of the fulcrum (First Order). The effort is applied very near to the fulcrum and so the mechanical advantage is small. The speed, however, is correspondingly great.

MOVEMENTS AT THE ANKLE JOINT

There are many movements at the ankle joint. We shall consider only two. Fig. 150 illustrates the mechanism of the foot. Examine the figure. What happens when the body is raised on the toes? The weight of the whole body falls on the foot at the ankle. The fulcrum is at the toes. The effort is supplied by the calf muscles, which are attached to the heel. The load is therefore between the fulcrum and the effort. There is a slight mechanical advantage. Negroes have long heels and correspondingly less-developed calf muscles because the lengthening of the heel gives an increase in mechanical advantage.
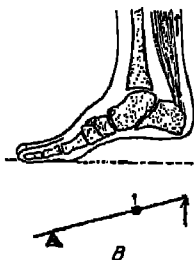
FIG. 150. The human foot.

Try to find out the relative positions of the load, effort, and fulcrum when the foot is bent as in raising the toes.

COUPLES

When two equal and parallel forces act on a body in opposite directions they tend to rotate the body. The two forces constitute a *couple*. We apply a couple when we turn a door knob to open a door (Fig. 151). The force F acting at A tends to rotate the handle clockwise about O, the moment of the force being $F \times (AO)$. Similarly the equal force F acting at B tends to rotate the handle in the same direction, the moment of the force being $F \times (BO)$. The moment of the couple is the sum of these two moments, namely, $F \times (AO) + F \times (BO)$, i.e. $F \times (AO + BO)$, i.e. $F \times (AB)$, i.e.
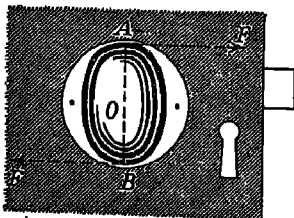
FIG. 151. A door-knob.

Moment of a couple = Either force × Arm of the couple.

We apply a couple when we turn on a tap or use a screw-

driver. Fig. 152 shows how the couple applied rotates the screw-
driver about its axis *XY*.

WORK

In ordinary language the term work has many meanings. To
some it means something done for
wages received. To others it is a hard
task which would not be performed
unless some advantage is obtained.
What one person calls work another
calls play. Why is it work to do a hard
sum for homework and play to solve
a difficult crossword puzzle? To the
student of science neither of these
performances is work at all. An
architect may be 'working' very hard
in his study, planning a building and
putting his ideas on paper, but in the
scientific meaning of the term he is
not doing anything like the amount of
work that his little daughter is doing
skipping in the garden. Yet the
daughter, if asked, would say that she was playing. She
would certainly be surprised if she were told she was working.

FIG. 152. A screw-driver.

When a force acts on a body and causes it to move, work is
said to be done. A horse which pulls a cart from one place to
another does work. You do work when you lift a parcel from
the floor and place it on a shelf. A man climbing a ladder does
work because he raises himself from the floor to the top of the
ladder. A girl skipping keeps raising herself a few inches from
the ground, and every time she does this she does work, but, of
course, in this case the net amount of work done is nought.

A weight held in the hand at arm's length and kept still
needs effort but no work is done in the scientific meaning of the
term. If you lift a pound weight from the floor to a height of
1 ft. you have done what is called a *ft.-lb.* of work. If the
weight is raised 2 ft., 2 ft.-lb. of work have been done. The
same work is done when a 2 lb. weight is raised 1 ft. The work
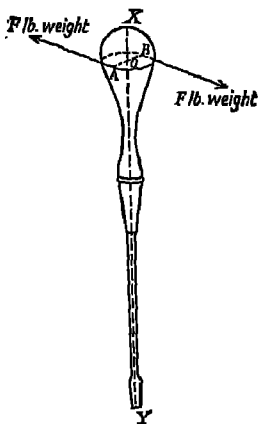
done against gravity is measured in ft.-lb. and is the product of the weight and the vertical distance through which it is raised. Generally

Work done = Force × Distance travelled in the direction of the force.

The ft.-lb. is called the *gravitational unit of work*. It is the work done in moving a mass of 1 lb. vertically through 1 ft. In the C.G.S. system the gm.-cm. is the unit.

Work may be useful or not. You may do work by jumping from the floor on to a chair twenty times. Apart from the exercise, it would not do you or anyone else any good, but it would be work. Suppose a workman on the roof of a house requires some slates which are on the ground. A labourer carries them up. The work done is:

(Wt. of man + Wt. of slates) × (Vertical distance from ground to the roof).

Only part of this is useful work. That done in raising the labourer from the ground to the roof is useless. If the slates had been raised by a pulley, nearly all the work would have been useful. This is one of the advantages of using pulleys.

THE WINDLASS

We have seen that the ordinary lever is not a suitable machine for raising loads great distances. A machine which embodies the principle of the lever may be used for this purpose. The windlass (Fig. 153) is really a continuous lever. The illustration shows a windlass used for raising water from a well. Round a cylinder is wound a rope, to one end of which is attached a bucket. The cylinder is turned by means of a handle fastened to the axis of the cylinder. The effort is applied at the end of the arm of the handle. The fulcrum is the pivot in line with the axis of the cylinder. The load is the bucket and what it contains. Consider Fig. 153. Let L be the load, E the effort, R the length of the arm of the handle, and r the radius of the cylinder. From the principle of the lever:

$$L \times r = E \times R.$$

Therefore, $L/E = R/r.$

That is, the mechanical advantage is R/r.

By making the length of the arm of the handle large compared with the radius of the cylinder the effort will be small compared with the load. There is, of course, a practical limit to the size of the arm. If it were very large no normal person would be tall enough to turn it. On the other hand there is a practical limit to the smallness of the radius of the cylinder.
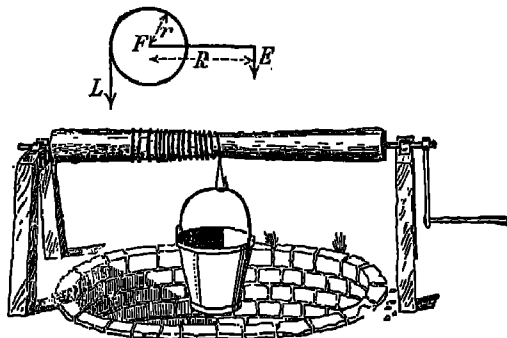
FIG. 159.

For every turn of the cylinder the bucket is raised a distance equal to the circumference of the cylinder. If the radius is very small then each turn of the handle raises the bucket very little. A compromise has to be made. By making R/r very large the effort is small but many turns have to be made to raise the bucket to the top. If R/r is not so large, the effort is larger but the bucket is raised to the top with fewer turns of the handle.

THE PRINCIPLE OF WORK

Let us consider the work done by the windlass in raising a bucket of water from the bottom of the well. The raising of the water is useful work, but other work has to be done by the machine. The rope has to be raised, and work has also been done against the friction of the moving parts. If we can neglect all friction and also the work done in raising the moving

parts of the machine, then the work done by the machine is measured by the product of the load and the distance the load is raised.

$$\text{Work done by the machine} = L \times d \text{ ft.-lb,}$$

where L is the load in pounds-weight and d is the distance the load is raised in feet. But the distance travelled by the load during one turn is $2\pi r$ ft. Hence, the work done on the load, that is, the work got out of the machine, is $2\pi r L$ ft.-lb. During one turn the effort (E) moves $2\pi R$ ft. Hence the work done by the effort is $2\pi R E$ ft.-lb. Now assume that the work done on the machine by the effort is equal to the work got out of the machine (neglecting friction).

Then, $2\pi R E = 2\pi r L.$

That is, $L/E = R/r.$

That is, the mechanical advantage is R/r.

We proved this before by using the principle of the lever. This time we proved it by assuming that the work done by the machine against the load is equal to the work done on the machine by the effort. This assumption is known as the *principle of work*. It states that all the work put into a machine is delivered by the machine in some form or other. In our example we assumed that all the work put into the machine was used in doing useful work. This is never the case in practice. The principle of work is often used in solving problems on machines.

EFFICIENCY AND VELOCITY RATIO

The efficiency of a machine is the ratio of the work got out to the work put in.

$$\text{Efficiency} = \frac{\text{Output}}{\text{Input}}.$$

If we could imagine a perfect machine in which none of the work put in was lost against friction, then, without operating the machine, it would be possible to calculate its mechanical advantage if we knew the designs of the various parts. This we have done when calculating the mechanical advantage of the

machines we have so far considered. In practice, where a machine cannot be considered perfect, the mechanical advantage must be obtained by experiment.

A very useful ratio, not affected by the work done in overcoming friction, is the *velocity ratio*.

$$\text{V.R.} = \frac{\text{Distance moved by the effort in a given time}}{\text{Distance moved by the load in the same time}}$$

$$= \frac{\text{Distance moved by the effort}}{\text{Corresponding distance moved by the load}}.$$

This can be calculated from a diagram of the machine. In the case of the windlass we found, by assuming it to be a perfect machine, that the mechanical advantage was R/r. The velocity ratio is $2\pi R / 2\pi r = R/r$. This is the same as the mechanical advantage because we considered the machine perfect. In a perfect machine the mechanical advantage equals the velocity ratio.

In any machine,

$$\text{Efficiency} = \frac{\text{Work got out}}{\text{Work put in}}$$

$$= \frac{\text{Load} \times \text{Distance moved by load}}{\text{Effort} \times \text{Distance moved by effort}}$$

$$= \frac{\text{Load}}{\text{Effort}} \times \frac{\text{Distance moved by load}}{\text{Distance moved by effort}}$$

$$= \frac{\text{Mechanical advantage}}{\text{Velocity ratio}}.$$

The efficiency of a perfect machine is, of course, 1, or 100 per cent., and the mechanical advantage is equal to the velocity ratio. In practice, the mechanical advantage is less than the velocity ratio.

In the case of a windlass, a bucket of water has to be raised to the top of a well. This represents so much work. This amount of work at least must be put into the machine. In practice more will be necessary on account of friction, &c. No

matter what the velocity ratio is, the work has to be done. When it is arranged so that the velocity ratio is large, the effort required will be small, but the effort will have to travel a long way before the load reaches the top. In other words, with a large velocity ratio the work is done with a small effort but at a slow speed; with a small velocity ratio the effort required is larger, but the work is done more quickly. It amounts to this: to get a definite amount of work out of a machine at least that much must be put in. It may be put in slowly or quickly. If it is put in slowly the velocity ratio will be large; if put in quickly, the velocity ratio will be small. What is gained in speed is lost in effort, and what is gained in effort is lost in speed.

THE INCLINED PLANE

You know it is easier to push a barrow up a gentle slope than up a steep one. Suppose you wish to lift a barrow with a load of bricks from the ground to the top of an embankment 5 ft. high. How could you do it? No matter what method you adopt the same amount of work has to be done. If the weight of the barrow and bricks combined is 200 lb., then 1,000 ft.-lb. of work has to be done in raising the load 5 ft. Suppose you place a strong plank leading from the ground to the top of the embankment and wheel the barrow up it. If the plank is 10 ft. long then you have to push 10 ft. before the load is raised 5 ft. vertically. If the plank is 20 ft. long you will have to push the load 20 ft. before the load is raised 5 ft. vertically, but in this case you have not to push so hard.

$$\text{V.R.} = \frac{\text{Distance through which effort moves}}{\text{Distance moved by load}}$$

$$= \frac{20}{5} = 4.$$

Assuming that all the work put in is given out, the mechanical advantage is equal to the velocity ratio, namely, 4. A load of 200 lb. weight can be thus raised by an effort of 50 lb. weight.

If the 10-ft. plank is used then the effort required is 100 lb. weight, but this has to be applied only over half the distance. The longer the slope the smaller is the effort required, but the effort must be applied through a longer distance.

In Fig. 155 L is the load, E the effort applied parallel to the
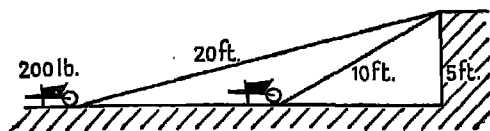
FIG. 154. The inclined plane.
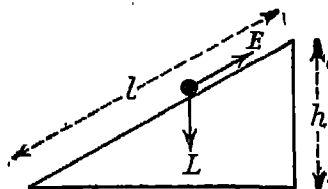
FIG. 155.

slope which is l ft. long, h is the vertical height to which the load is raised.

$$\text{V.R.} = \frac{l}{h} = \frac{\text{Length of slope}}{\text{Vertical height}}.$$

$$\text{Efficiency} = \frac{\text{M.A.}}{\text{V.R.}} = \frac{L}{E} \times \frac{\text{Perp. height}}{\text{Length of slope}}.$$

For a given height the velocity ratio is increased by lengthening the slope. Our everyday experience tells us that this is true.

The inclined plane is of great practical value and is often used during building operations. It is thought that the great Pyramids of Egypt were built by men pushing the great slabs of stone on rollers up inclined planes. In permanent structures also the inclined plane is of great use. A staircase is an inclined plane which has notches in it for your feet. A spiral staircase is a device for employing the principle of the inclined

plane in a small area. No doubt you have noticed the tendency for horses pulling heavy loads up difficult hills to 'tack' across the road. This is encouraged by the driver when road condi-tions permit. The horse is increasing the velocity ratio.

THE WEDGE

The wedge is another simple machine. Its principle is employed in the hatchet, knife, and chisel. Apart from these instruments the wedge itself is often used for splitting wood, forcing two bodies apart, and sometimes for raising heavy bodies a little distance from the ground.

Expt. 227. Examine a hatchet, knife, and chisel and use them for their proper purposes noting how these purposes are effected. Obtain a steel wedge commonly used for splitting wood and, while using it, try to discover how it works. Where do you apply the force? What resists the force you apply? In what direction is this resistance? Place a wooden wedge under the leg of a heavy table and drive the wedge gently between the floor and the leg. Repeat the above experiments using thick and thin wedges. Do you notice any difference?

Most, but by no means all, wedges are so shaped that the end faces are isosceles triangles (Fig. 156a). Consider such a wedge, ABC (Fig. 156b), being driven between a heavy object, P, and the floor, Q. When a force, E, is applied perpendicular to the base, AB, of the wedge, it tends to push P and Q apart and, of course, P and Q tend to resist the insertion of the wedge. If the wedge is smooth and *if we neglect friction* then it can be proved that:

(1) the reactions, G, F, by which P and Q resist the insertion of the wedge, act at right angles to the sides of the isosceles triangle as shown in the figure;

(2) the reactions F and G are equal and each is equal to the effort E multiplied by the ratio of the side to the base; or, $F = G = E \times (BC/AB)$.

For example: If $AB = 1$ in. and $BC = 5$ in., then $F = 5E$; that is, the force tending to push the objects apart is five times that tending to push the wedge in. If the base of the wedge is made very small and the length as large as is practicable then a very small force applied to the base of the wedge can be made to overcome a very large resistance.

Of course, in practice there is much friction. Indeed, if there were no friction, the wedge would slip out when the effort was removed. There must therefore be enough friction to keep the wedge in when the effort is removed. The facts we have stated above apply only to an imaginary frictionless wedge. In practice much of the effort applied to drive the wedge in is used in

(a)   (b)
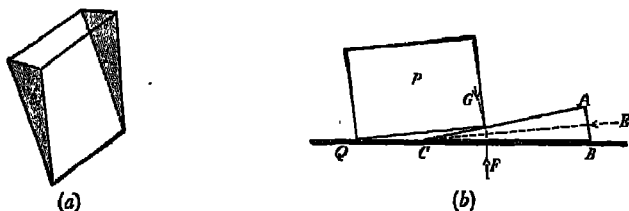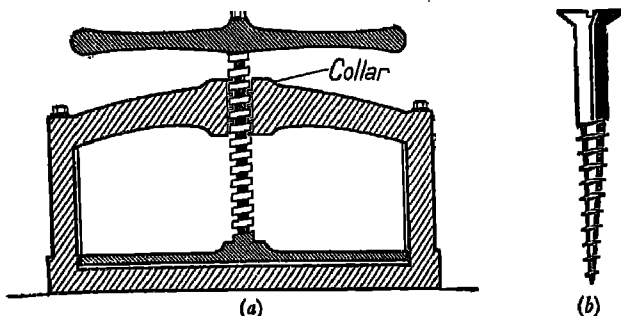
FIG. 156.

FIG. 157. (a) screw-press; (b) screw for wood.

overcoming friction, but it remains true that the smaller the base of the wedge compared with the length, the smaller is the effort required to drive the wedge in. You can very easily test this practically in splitting wood by using a short thick wedge and then a long thin one.

THE SCREW

When we think of screws we usually think of those which are screwed into wood by means of a screwdriver, but there are other common kinds such as cylindrical screws which work in

a' fixed collar or nut. These may be seen in such appliances as the vice, screw-press, and screw-jack.

Examine as many appliances involving screws as you can. Note how they work and pay particular attention to the screws (Fig. 157). The common screw which goes into wood is conical with a spiral thread. Others, working in a collar, are cylindrical and they also have a spiral thread. In most cases the thread is a prominence on the cylinder but there are some threads which are grooves in the cylinder. The threads which are prominences fit into corresponding grooves in the collar and vice versa.
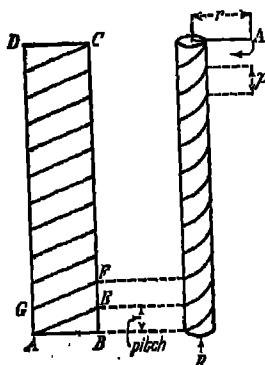
Take a wooden cylinder any convenient size (or thick candle will do) and measure its circumference with a piece of string. Cut out a rectangle of paper as shown in Fig. 158, making AB the length of the circumference and BC any length. Mark off equal intervals along BC and equal intervals of the same length along AD. Join EA and FG and so on. Now wrap the rectangle of paper about the cylinder. You will find that the inclined planes AE, GF, &c., form a continuous spiral thread round the cylinder showing that a screw thread is really a spiral inclined plane.

Fig. 158.

Note that you could have made the intervals BE, EF, &c., much smaller (or larger) as long as they were all equal. BE, that is, the distance between two consecutive threads measured parallel to the axis of the screw, is called the *pitch* of the screw. If there are 10 threads to the inch the pitch is said to be 1/10 in. Examine as many different screws as you can and, by measurement with a ruler and counting, find the pitch of each.

Give one complete rotation to the handle of a screw-press or one complete turn to a screwdriver driving a screw into wood, and note how far the end of the screw travels. Clearly it must travel the length of its pitch.

It should now be clear how to find the velocity ratio of a screw. Consider Fig. 158. Let r be the radius of the arm at the end of which, A, the effort, E, is applied. In one rotation of the screw the effort travels $2\pi r$. This causes the screw to move against the resistance, R, a distance p equal to the pitch of the screw. Hence

$$V.R. = \frac{2\pi r}{p}, \quad \text{or} \quad \frac{\text{Circumference of circle described by arm}}{\text{Pitch}}.$$

A screw with a very small pitch has a very large velocity ratio and thus a small force is transformed into a large one, but, as a consequence, rapid motion is transformed into slow motion. Theoretically we could have the pitch as small as we like but there are practical difficulties. The prominences of the thread would not be strong enough to stand the strain and also many rotations would be required to force the screw only a very small distance.

FIG. 159. A grinding wheel.

If the screw is rotated by a screwdriver to which a couple is applied (p. 340) the calculation of the velocity ratio is just the same, as the couple may be considered as being twice the effort on one radius.

As we have seen, common applications of the principle of the screw are the screw-press, the vice, and the screw-jack. The latter usually employs, in addition, the principle of geared wheels and we consider this application on page 354.
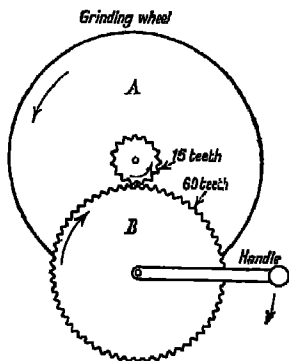
GEARED WHEELS

It is often necessary to transfer power from one shaft to another. It has to be done in factories, bicycles, motor-cars, watches, and almost every piece of machinery. When the shafts are some distance apart, as in factories, belts or ropes are used to transmit the power from one shaft to another, but when the shafts are close together, as in a watch, toothed wheels are used.

Examine some simple machine employing toothed wheels such as a grinding wheel (Fig. 159). Note what happens when you turn the handle of the driving wheel. The grinding wheel, A, revolves much

faster than the driving wheel, *B*. How much faster? Count the number of teeth in each wheel and divide the smaller number into the larger. This gives the number of revolutions of the smaller wheel for one revolution of the driving wheel. If, for example, there are 60 teeth in the driving wheel and 15 in the grinding wheel, the latter revolves 4 times (60/15) as fast as the driving wheel.

Note the force you must apply to the handle of the driving wheel to start it. Now try to start the grinding wheel by pushing it with your finger and thus move the driving wheel. Note that it is very easy to move the grinding wheel in this manner but not so easy to move
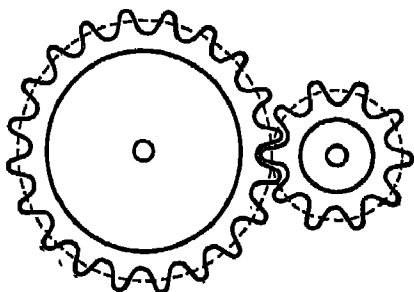
Fig. 160. Two cogwheels in mesh.

the driving wheel. In the example we have considered the grinding wheel is moved by only $\frac{1}{4}$ (i.e. 15/60) of the force necessary to move the other wheel.

Note also that the two wheels revolve in opposite directions, *A* counter-clockwise and *B* clockwise.

Two geared wheels cannot be in mesh unless the teeth fit properly. Consider Fig. 160, which represents two cogwheels in mesh. The circumferences of the wheels are represented by broken lines. The teeth project beyond the circumference and also are indented into it. The distance, measured along the circumference, between two adjacent teeth is called the pitch of the teeth. Two cogwheels cannot be in mesh unless they have teeth of the same pitch. The diameter of the wheels does not affect the pitch but it does affect the number of revolutions. Number of revolutions of the smaller cog for one revolution of the larger

$$= \frac{\text{Number of teeth of the larger cog}}{\text{Number of teeth of the smaller cog}},$$

$$= \frac{\text{Circumference of the larger cog}}{\text{Circumference of the smaller cog}},$$

$$= \frac{2\pi R}{2\pi r}, \text{ where } R \text{ is the radius of the larger cog and}$$
$$r \text{ the radius of the smaller cog,}$$

$$= \frac{R}{r}.$$

It is often necessary to have a combination of more than two
geared wheels in mesh as in clocks
and watches, for here different
speeds of revolution are required
for different wheels. In the case
of some cars, however, a combina-
tion of more than two geared
wheels is used to produce revolu-
tions in the same direction as the
driving wheel. This arrangement is known as a train of gear
wheels.
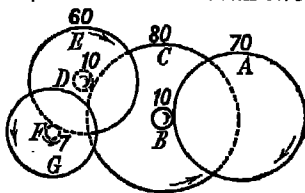
FIG. 161. A train of wheels.

Examine the train of wheels in an old discarded watch or clock.
Take the watch to pieces and find out how the cogs work. Count the
number of teeth in the wheels and calculate the relative speeds of
revolution of the wheels. If you cannot procure an old watch then
try to work from Fig. 161, but it is much more interesting to use the
actual wheels. Examine as many different machines as you can and
try to understand how the train of wheels works.

In Fig. 161 the four wheels *A*, *C*, *E*, *G* turn when *A* is turned.
This is effected by means of the toothed wheels and the cogs
B, *D*, *F*. The number of teeth in each is indicated in the figure.

    1 rev. of *A* causes 7 (i.e. 70/10) revs. of cog *B* and therefore
       of *C*.

    7 revs. of *C* cause 56 (7 × 80/10) revs. of cog *D* and therefore
       of *E*.

    56 revs. of *E* cause 480 (56 × 60/7) revs. of cog *F* and there-
       fore of *G*.

Or, if *A* revolves in a clockwise direction then:

  G revolves (counter-clockwise) 480 times as fast as *A*

  E revolves (clockwise) 56 times as fast as *A*.

  C revolves (counter-clockwise) 7 times as fast as *A*.

  Geared wheels may, of course, be used for reducing, as well as for increasing, the speed of revolution.

THE SCREW-JACK

  The type of screw-jack illustrated in Fig. 162 is in general use among motorists for raising one side of the car while a
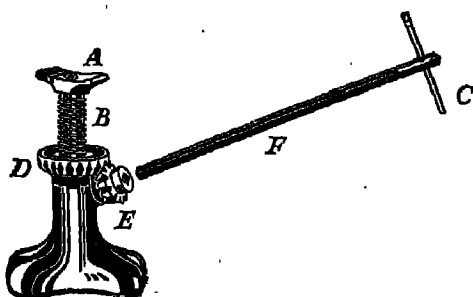
Fig. 162. The screw-jack.

wheel is being changed. The load rests on the cap *A* which is free to turn in the head of a screw *B*. The screw fits into a toothed wheel *D*. When this wheel is turned the screw is turned and the load raised. The teeth of *D* engage in the teeth of another wheel *E* which is turned by the cross-bar *G* rotating the rod *F*. The rod *F* and the wheel *E* are inclined at an angle to the horizontal so that the operator can conveniently use the jack.

  Let the length of the cross-bar be *x* in., *n* the number of teeth in wheel *E*, *N* the number of teeth in wheel *D*, and the pitch *v* in. Of course, *y* will be a fraction of an inch. The effort is applied to the ends of the cross-bar. What is the velocity ratio? During one rotation of the bar *F* the effort moves πx in. This will cause one complete revolution of the wheel *E*, but

not one revolution of the wheel D because the number of teeth is different. The wheel D will have rotated n/N revolutions and the load will have moved $n/N \times y$ in.

$$V.R. = \frac{\text{Distance moved by effort}}{\text{Distance moved by load}}$$

$$= \frac{\pi x}{\frac{n}{N} \times y} = \frac{\pi x N}{ny}.$$
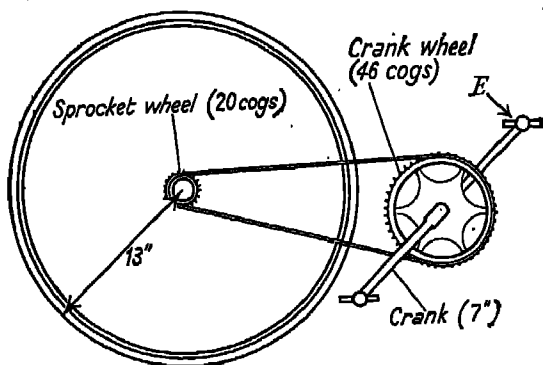
FIG. 163. A bicycle.

Do not try to remember this expression for the velocity ratio, but try to understand the method. Substitute suitable values for the symbols and work out the velocity ratio for a particular screw-jack.

THE BICYCLE

The force is applied to the pedals which are attached by means of cranks to the toothed crank wheel (Fig. 163). An endless chain connects the crank wheel to the toothed sprocket wheel which is attached to the hub of the back wheel. The number of teeth on the sprocket wheel is smaller than that on the crank wheel, so that for one revolution of the crank wheel the back wheel revolves more than once.

Expt. 228. To find the velocity ratio of a bicycle. Measure the diameter of the back wheel and the length of the crank connecting the pedal to the crank wheel. Count the number of cogs (teeth) on the crank and sprocket wheels. In a particular case the following results were obtained:

Length of crank   .   .   .   .   .   .   7 in.
Diameter of back wheel   .   .   .   .   . 26 ,,
No. of cogs on crank wheel   46
No. of cogs on sprocket wheel   20

In one revolution of the crank wheel the effort moves $2\pi \times 7$ in. During this time, the sprocket wheel, and therefore the back wheel, does 46/20 revolutions. Hence the distance travelled by the bicycle during one revolution of the crank wheel is

$$\frac{\pi \times 26 \times 46}{20} \text{ in.}$$

This is also the distance moved by the load against all the resistances.

$$\text{Therefore V.R.} = \frac{2\pi \times 7 \times 20}{\pi \times 26 \times 46} = \frac{70}{299} = \frac{7}{30} \text{ approx.}$$

The mechanical advantage cannot be calculated from a diagram, but it is clear that the mechanical advantage of a bicycle is much less than 1. The force required on the pedal is much more than that on the back wheel. The object is not to overcome a large resistance with a small effort, but to overcome a small resistance with a large effort so that speed will be obtained. When bicycling up a steep hill it is found that this arrangement is not very good. The effort required is too large. It would be an advantage if the effort could be reduced and the speed reduced. This is possible on some bicycles by changing the gear.

On the old 'penny farthing' bicycle the cranks operated the large wheel directly. One revolution of the pedals caused one revolution of the large wheel. The larger the diameter of this wheel the farther the machine would travel for one revolution of the pedals. If the gear of a modern bicycle is 60, then for one revolution of the pedals it will travel the same distance as if the pedals were attached directly, as in the old bicycle, to a wheel 60 in. in diameter. Hence

$$\text{Gear} = \frac{N}{n} \times \text{Diameter of the back wheel,}$$

where N is the number of cogs on the crank wheel and n the number on the sprocket wheel. The gear of a bicycle thus depends on the size of the crank wheel and sprocket wheel and on the diameter of the back wheel.

It is clear that if the gear is increased the velocity ratio is decreased. A low gear is employed when ascending a hill so that the velocity ratio will be greater. The speed of the machine will be reduced but the effort required will also be reduced.
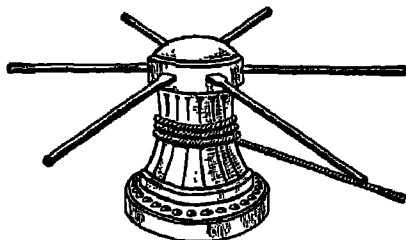
FIG. 164. A capstan.

The effort, of course, has to travel a correspondingly greater distance.

THE CAPSTAN

The capstan is a similar machine to the windlass, and uses the same principle (Fig. 164). It is used for raising the anchor of the ship, or for pulling the ship to the quay side. The effort is applied by men to the arms projecting from the barrel. In most modern ships the effort is supplied by an engine, but before the use of steam hand capstans were universal on ships. The velocity ratio is the same as in the windlass. The practical limit to the length of the arms is the width of the ship at the part where the capstan is fixed. The arms are removed when the capstan is not in use.

PULLEYS

A pulley (Fig. 165) is a grooved wheel turning on an axle so arranged that a rope (or chain or string) can travel in the groove. The framework enclosing the wheel is called the block. This may be fixed or movable. You should watch workmen

using pulleys and make some for yourself. The wheels are easily obtained and string may be used for the rope. Study this paragraph and then test the results as far as possible with model pulleys.

A *fixed pulley* is arranged as in Fig. 166. Clearly, when the
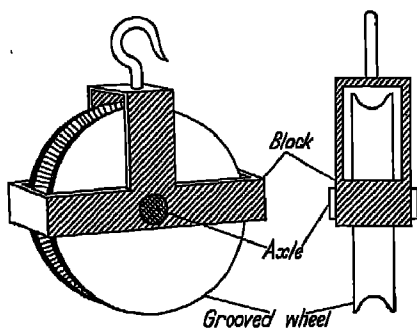
FIG. 165. A pulley.
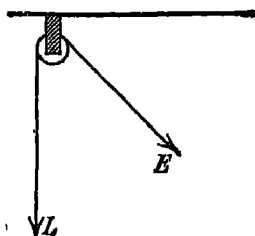
FIG. 166. A fixed pulley.

effort moves through 1 ft., the load also moves through 1 ft. The velocity ratio is therefore 1. Let E be the effort and L the load. Neglect friction and the weight of the rope.

$$E \times \text{Radius of pulley} = L \times \text{Radius of pulley}.$$

Hence, $E = L.$

This result is also clear from the principle of work. The

load has to travel the same distance as the effort, d ft. If we consider the machine perfect so that all the work put in is got out, then:

Work done by the effort = Work done on the load,

$$E \times d = L \times d,$$

$$E = L.$$

The mechanical advantage is therefore 1. This is equal to the velocity ratio as in all perfect machines.

Although the effort is the same as the load, this pulley is nevertheless convenient in practice, because the effort can be applied in a different direction. It is easier for a man to pull downwards than to pull upwards. By using this arrangement a horse can be employed to pull a load to the top of a building.
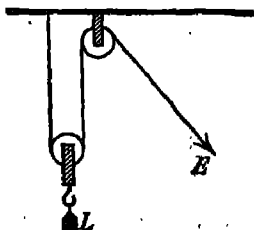
FIG. 167. Single movable pulley.

A *single movable pulley* is commonly used on cranes and is arranged as in Fig. 167. When the effort moves 2 ft. the load moves 1 ft. The velocity ratio is therefore 2. Apply the principle of work neglecting friction and the weight of the rope. Let L be the combined weight of the load and the block, E the effort, and h the height the load is raised.

$$L \times h = E \times 2h,$$

$$L/E = 2.$$
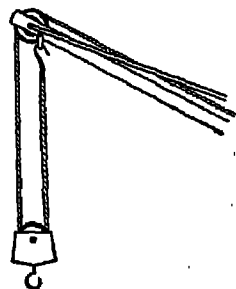
FIG. 168. A crane with a single movable pulley.

The mechanical advantage is 2. This is the same as the velocity ratio, for we have considered the machine perfect.

There are many types of *single-rope pulleys*. One of the most useful is represented in Fig. 169. There are two blocks, one fixed, and one movable. Each block contains the same

number of wheels, and the wheels of each block work on the same axle.

Let n be the number of parts of the rope supporting the lower block. (This will be the same as the total number of pulleys in the two blocks.) When the effort moves n ft. the load moves 1 ft. The velocity ratio is therefore n, that is, the same as the number of pulleys in the two blocks. Apply the principle of work using the same notation as before, and neglecting friction. Then

$$L \times h = E \times nh,$$

$$\frac{L}{E} = n.$$

Fig. 169. A single-rope pulley with two wheels in each block.

That is, the mechanical advantage is equal to the total number of pulleys, and the machine being considered perfect, is equal to the velocity ratio.

FRICTION

We have already said something about friction (p. 309) and have noted its importance in the construction of machines. Friction may be useful, as in brakes and the transmission of power by belts. It is often a disadvantage, and machines are designed to work with as little friction as possible.

Fig. 170. A single-rope pulley (two wheels in each block) used as a crane.

Friction always occurs wherever two surfaces are sliding over each other. No bodies are perfectly smooth, and there is always some force between two bodies in contact which prevents one from sliding over the other.

Expt. 229. Attach a length of string to a hook screwed into a wooden block resting on a horizontal table. There is a force due to the weight of the block, acting downwards, that is, normal to the table. Now pull the string gently in a horizontal direction. The block does not move, but you can feel the resistance to your pull. Since the block does not move, the resist-
ance, or friction, must be equal to the force of your pull. Now stop pulling. The block remains at rest. Friction cannot now be acting. Pull again with gradu-ally increasing force. So long as the block remains at rest the force of friction increases to balance the force exerted by the string. It is, of course, in the opposite direction. As the pull of the string increases there comes a time when the block begins to move along the surface of the table. The force of friction is not sufficient to over-come the pull of the string, but

FIG. 171.

friction is still acting in the opposite direction to the pull of the string tending to prevent the block from moving.
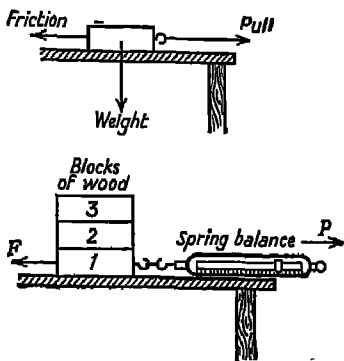
The friction force when the block is just on the point of sliding is called *limiting friction*. It is, of course, equal to the force which will just start the block moving. The force acting against the pull of the string when the block is moving is called *sliding friction*. It is equal to the force which will just keep the block moving at a steady speed.

Expt. 230. Attach a delicate spring balance (or dynamometer) by means of a string to a hook in a block of wood having rectangular faces. Let the adjacent faces be unequal in area. Arrange the apparatus on a table as shown in Fig. 171. Pull the dynamometer gently and note the reading when the block just begins to move. Continue pulling, and note the reading as the block slides slowly along the surface of the table. Put another similar block on top of the first and repeat the readings. Repeat, with a third block on top of the second. Now turn the block on another face and arrange the hook so that the dynamometer is pulling in the same direction as in the previous experiments. Repeat all the previous readings. Repeat the experiments with blocks made of glass and metal. Tabulate your readings as shown on the next page.

Wt. of the block(s) in gm. W	Force causing the block(s) to start moving (gm.) F	Force which just keeps the block(s) moving (gm.) f	$\dfrac{F}{W}$	$\dfrac{f}{W}$

Your results should show:

1. The greater the weight the greater is the friction, or, the force of friction varies as the normal reaction (the force acting downwards due to the weight of the blocks).
2. Friction acts in the opposite direction to the motion.
3. Friction is independent of the areas of the surfaces in contact.
4. Sliding friction is less than limiting friction.

The ratio F/W is the *coefficient of statical friction*, f/W the *coefficient of sliding friction*.

Consider an engine pulling a train. As the engine starts it has to overcome the friction at the axles and at all the moving parts. This requires a greater force than must be exerted to overcome the friction after the train has started to move.

ROLLING FRICTION

Drag your bicycle sideways across the road and note how much harder it is than wheeling it in the ordinary way. Wheels were invented in prehistoric times by someone who wished to reduce the friction present when there was a sliding contact between the load and the ground. Rolling friction is less than sliding friction. It is caused by the body, owing to its weight, sinking a little into the other surface and it has to be lifted out of the depression thus made. The wheels of an engine, for example, cause a little depression in the rails. As the engine moves it has to be lifted on to the next portion of the line and then sinks again. You will get a good idea of this if you ride your bicycle first on a smooth road and then on a sandy path, and note the difference. The weight of an engine is so great

that it would sink into the track if there were no rails. These reduce the friction.

Rolling contact is also used to reduce the friction in machines when possible. Ball-bearings are common. Examine your bicycle and find out where ball-bearings are employed. Support the bicycle off the ground and turn the wheel by giving it one push. Note how many revolutions are made before it comes to rest. Tighten the cone of the axle and try again. The number of revolutions is smaller. The friction has increased. Tightening the cone is equivalent to increasing the load, and rolling friction obeys the same law as sliding friction in that it varies with the load.

LUBRICATION

Friction is considerably reduced in machinery by the use of lubricants. These are substances placed between the sliding or rolling surfaces, the effect being to reduce the friction. Oil and grease are well-known lubricants. For light machinery like watches a very thin oil is used, but for heavier machinery thicker oil and grease are used. The thick oil remains where it is required whereas the thin oil would be forced out. The temperature at which the machinery is working influences the type of oil used. Graphite (p. 179) is used on chains and cogwheels. Sometimes the graphite is mixed with grease. Graphite without the grease is a good lubricant for wood. Soap is also used for this purpose. When a ship is being launched the slip-ways are liberally spread with soft soap. Some parts of a machine do not require lubricating. They depend on friction for correct working. Grease on belts causes slipping which defeats the object of the belt. When oiling a bicycle it is important to keep oil from dripping on the wheel rims which would cause the brake-blocks to slip and thus reduce the braking power.

It is easy to see that lubrication does reduce friction. A bicycle runs much more easily if, after long neglect, it is oiled. A well-oiled lock is different from the usual neglected one. As friction also causes wear, reducing friction by lubrication also reduces wear. Well-oiled machinery lasts longer. (See p. 487 for the lubrication of an aero-engine.)

HEAT AND FRICTION

Expt. 231. Rub a piece of wood vigorously with sand-paper. You can feel that the sand-paper and wood become hot. Feel the brake and tyres of a motor-car when it stops after coming down a steep and long hill with the brakes on. They are hot. Sharpen a tool on a grindstone. Water (or oil) is used to keep the tool from becoming too hot and so spoiling the temper of the metal. Tighten the cone of the axle of a bicycle and then turn the wheel vigorously for a minute. Feel the axle. It is hot.

To overcome the force of friction work has to be done. When you saw through a piece of wood, the saw travels a certain distance against the resistance of the wood. The work done appears as heat. This is why some of the work put into a machine cannot appear as useful work. Some is used in overcoming friction and the machine becomes hotter. The greatest heat appears in those parts of the machine where most friction occurs. As metals are damaged when they become too hot, this is another reason for reducing friction in machines.

ENERGY

The capacity for doing work is called *energy*. We ourselves (and all living things) have some capacity for doing work and so we are said to possess energy. People who are always doing work willingly are called energetic, but this is only a general descriptive term, and is not used scientifically. The amount of energy possessed by a body is the amount of work it is capable of performing. The unit of energy is therefore the same as the unit of work. There are other forms of energy, e.g. heat, light, sound, electrical energy, and chemical energy. We shall see later (Ch. XLIII) that one kind of energy can be transformed into another kind. We have already seen that mechanical energy is commonly transformed into heat when work is done against friction.

POTENTIAL ENERGY

Have you ever watched workmen driving great piles into the earth? A heavy weight is wound to the top of a framework and then allowed to fall on to the top of the pile. As it does so, the pile is driven a little way into the ground. The process is repeated until the pile is driven as far as is required. When the weight is at rest at the top of the framework, it is capable of doing work by virtue of the position it holds. It is said to have *potential energy*. Work is 'stored up' in the weight. If the weight is 100 lb., and the distance of the top of the framework from the top of the pile is 20 ft., the weight possesses by the virtue of its position 2,000 ft.-lb. of energy, because it is capable of doing that much work. In other words, 2,000 ft.-lb. of work is stored in the weight. When the weight falls and reaches the top of the pile, it no longer possesses that energy. The stored-up work in the weight is expended in driving the pile a little way into the earth against friction. The water in a mill dam has potential energy because it is capable of doing work by turning the water-wheel that grinds the corn. The wound-up spring of a watch has potential energy. The work stored up in it is used to move the wheels and levers of the watch against friction. A grandfather clock is worked by a weight which is wound to the top. In that position it has potential energy, because in falling it can do work.

KINETIC ENERGY

A moving object is capable of doing work in coming to rest. Consider a bullet fired at a wooden target. The bullet penetrates the target. As it does so it loses speed and finally comes to rest. This capacity for doing work (in this case against the resistance of the wood due to friction) by virtue of its motion is called the *kinetic energy* of the body. Consider the heavy weight used in pile-driving. When it is at rest at the top of the framework it has potential energy, but when it starts to fall the potential energy is transformed into energy of motion, or kinetic energy. It is this form of energy that actually does the work of

driving the pile. A moving hammer head has kinetic energy because it is capable of doing work, on the head of the nail, in coming to rest. A stone thrown vertically upwards does work against gravity by virtue of its energy of motion. In doing this work it comes to rest for a moment in the air. At that moment it has no kinetic energy, because it is at rest, but it has potential energy because of its position. When the stone begins to fall the potential energy is transformed into kinetic energy again. A flowing stream has kinetic energy because, by virtue of its motion, it can do work, e.g. turn a water-wheel.
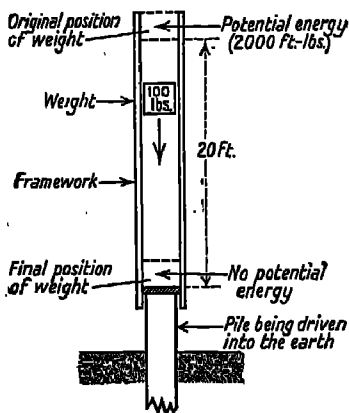
Original position of weight—— Potential energy (2000 ft.-lbs.)

Weight—— 100 lbs.

20 ft.

Framework——

Final position of weight—— No potential energy

Pile being driven into the earth

Fig. 172. Pile-driving.

The kinetic energy of a body is measured by the amount of work it is capable of doing in coming to rest.

THE CONSERVATION OF ENERGY

We have seen that potential energy can be transformed into kinetic energy and vice versa. When this occurs there is no loss of energy. The falling weight in pile-driving gains kinetic energy as it loses potential energy. But what happens to the energy when the weight strikes the top of the pile and comes to rest? The weight no longer possesses potential energy because it is the same height as the top of the pile. It no longer has kinetic energy because it is at rest. It is certain that all the mechanical energy has disappeared. If the weight and the top of the pile are examined it will be found that they are warmer than they were before the fall of the weight. The mechanical energy has been transformed into heat which is another form of energy. This is a very common transformation. When an iron chisel is hit several times with a hammer, both the chisel and the hammer head become hot. The mechanical energy is

transformed almost entirely into heat. We have seen that the brakes operating on a wheel become hot, and may become so hot that they burn away. The water at the bottom of a waterfall is slightly warmer than the same water when it was at the top. The water had potential energy before it fell; during its fall it had kinetic energy; this was transformed into heat at the bottom. (See Mechanical Equivalent of Heat, p. 483.)

Consider again the operation of pile-driving. Does the heat formed represent all the mechanical energy, or has some been lost or transformed into other forms of energy? When the weight falls on the top of the pile some noise is produced. Sound is a form of energy because it is capable of doing work (Ch. XXXV). Some of the mechanical energy is transformed into sound. It has been found by experience that whenever energy is transformed from one form to another no energy is lost. Whenever it has been possible to test the truth of this by experiment it has been done. Whenever men have worked on this assumption they have found their results to be true; and whenever they have assumed the opposite they have found their results to be false. For these reasons it is believed to be true. The fact is usually stated in the form: *energy cannot be created or destroyed.* This statement of what is believed to be a fact is called the LAW OF CONSERVATION OF ENERGY. Compare this law with the Law of Conservation of Mass (p. 161).

It is always taken for granted, because so far experience has shown it to be true, that whenever energy in one form disappears, an equal amount of energy either in the same form or in another appears. Consider what happens when a gun is fired. The chemical energy of the explosive charge disappears. The shell travels on its way, the gun jumps back a little, the shell and the gun become hot, there is a flash of light and a bang. In other words heat, light, sound, and kinetic energy are produced. According to the law of the conservation of energy, the sum of all these forms of energy is equal to the chemical energy of the explosive charge. Some of the energy is wasted. When a gun is fired, all that is required is that the shell should travel with all the energy possible. It is not required that the shell and the gun should become hot, or that the gun should jump back,

or that there should be noise and light. All these happenings
are disadvantages, but they are unavoidable. The energy
causing them is for all practical purposes lost, but this does not
affect the law of conservation of energy. Energy is continually
being lost in this way. To avoid confusion we say that energy
which is lost for all practical purposes is *dissipated*. This subject
is so important that it will often be referred to and in particular
we have devoted Chapter XLIII to its study.

To-day everyone is interested in *atomic energy*. An attempt
to study this subject in an elementary Science Book would be
out of place, but the subject is so important that we must, at
least, mention the relation of atomic energy to the Laws of
Conservation of Energy and Mass. These two laws are still true
so far as most chemical reactions and the working of machines
are concerned. When, however, atomic energy is released,
as in the atomic bomb, some matter is annihilated. It is now
possible to annihilate some forms of matter and this releases
energy, but only a definite amount of energy can be obtained
from a given mass of matter. If we wish to rewrite the Laws of
Conservation of Energy and Mass, so as to include the release of
atomic energy, we should write: The sum total of matter and
energy remains constant.

PERPETUAL MOTION

In considering the work done by machines, we assumed the
Principle of Work to be true. We now see that the principle
must be true if the Law of the Conservation of Energy is true.
All the work put into a machine appears in some form or
other. A part of it appears as useful work done outside the
machine. Another part is used in moving the movable parts
of the machine, and a third part is used in overcoming friction.
As it is impossible to have a frictionless machine, no machine
can give out as much work as is put into it.

There are still people who believe that it is possible to con-
struct a machine which will continue to go without further
effort after it has once been started. They dream of inventing
a motor-car which will go without petrol once it has been
started; or of making all the machines of a factory do work

without supplying any more fuel than is necessary to start them. They reason somewhat as follows: Drive a car along a road by means of a petrol engine. At the same time let the moving car drive a dynamo and use the electricity thus produced to drive the car. All the time the car is moving electricity is being produced and this can be used to move the car. This is a very crude example and not many people would believe it possible, but some are ready to believe that the moving car can produce some electricity which will help to keep the car moving and thus save some petrol. It was once suggested that a boat could be driven by a gas engine. The hydrogen used in the combustion would be oxidized to water and this could then be electrolysed and the hydrogen thus formed used again in the engine.

Before the principle of work was properly understood many workers spent much of their time trying to produce perpetual motion. Many ingenious devices were invented but they never got to the working-model stage. They looked all right on paper but they did not work. Perpetual motion by itself, even if realized, is of no use. What is wanted is a machine that will do external work as well. It is not difficult to construct machines where friction is so small that the machines will continue moving for a very long time. Eventually they stop because friction cannot be entirely eliminated. When any attempt is made to use the machine for external work it stops almost immediately.

The search for perpetual motion has not all been wasted effort. When the alchemists were searching for the Philosopher's Stone which would turn base metals into gold, they discovered many useful chemical facts although they never found what they were looking for. In the same way ingenious ways of doing work have been invented by the searchers for perpetual motion.

POWER

So far we have dealt with the amount of work a machine can do, but the *time* in which a machine can do a definite amount of work is also important. If work has to be done it is necessary to use a machine which can do the work in a reasonable time. Sometimes the time is all-important. Suppose water collects in a mine and has to be pumped to the surface. A pump must be used which can raise the water at least as quickly as it collects.

The rate at which an engine can do work is called its *power*. When the steam engine was coming into general use some unit became necessary by which the power of an engine could be measured. Previously horses had been employed to do the work, and most people had a general idea of how much work a horse could do in a certain time. This influenced JAMES WATT (1736–1819), the inventor of the modern steam engine, in his choice of unit. He experimented with a horse pulling a weight up a mine shaft, and thus found the work done by the horse per minute. This he called the *horse-power*. As a result of his tests he fixed this at 33,000 ft.-lb. of work per minute.

$$\text{One horse-power} = 33{,}000 \text{ ft.-lb. per minute,}$$
$$= 550 \text{ ft.-lb. per second.}$$

For a short period a horse is capable of working at a greater rate than this. Indeed, a man sometimes works for a short period at a rate of over 2 horse-power. Over a long period a man of ordinary vigour does not work at a rate of much more than one-tenth of a horse-power.

A 10 horse-power engine is capable of doing work at the rate of 5,500 ft.-lb. per second. This does not mean that it always works at this rate. It may, and generally does, work at a lower rate.

BRAKE HORSE-POWER

The horse-power of a machine may be calculated when certain facts about the machine are known, but a better estimate of the real capability of the machine is obtained by measuring the horse-power against the force of friction. The engine to be tested drives, through a belt, a shaft on which is a pulley wheel (Fig. 173). The pulley wheel has no groove, and over it passes another belt one end of which is connected to a spring balance which is fastened to a bench. The other end of the belt carries a scale pan into which weights may be placed. When the engine is working, the belt slips on the grooveless pulley wheel and weights are added to the pan until the belt keeps its position. The friction between the belt and the pulley wheel is measured by the difference between the weights and the reading of the spring balance. A counter is attached to the wheel so that the number of revolutions per minute can be found.
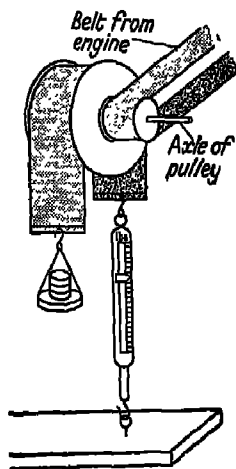
FIG. 173. To find the brake horse-power of an engine.

If the total weight in the pan be   .   .   .   . a lb.
 reading of the spring balance   .   .   .   . b ,,
 radius of the pulley wheel .   .   .   .   . r ft.
 number of revolutions per second   .   .   . n

Then frictional force $= (a-b)$ lb.

In one revolution the force acts through $2\pi r$ ft.

,, n revolutions ,, ,, ,, ,, $2\pi rn$,,

Rate of working $= 2\pi rn(a-b)$ ft.-lb. per sec.

Brake horse-power $= \dfrac{2\pi rn(a-b)}{550}$.

QUESTIONS

1. What is a machine? Illustrate your answer by reference to (a) a spade, and (b) a bicycle.

2. What do you understand by the term work? How is it measured? How much work does a 10-stone man do in carrying bricks weighing 25 lb. up a ladder to the roof of a house 20 ft. from the ground?

3. Explain carefully what is meant by the terms: (a) mechanical advantage, (b) velocity ratio, (c) efficiency of a machine. What is the relation between these three quantities?

4. Define energy and say how it is measured. What do you understand by the terms kinetic energy and potential energy?

Water evaporates from the ocean and is carried over land where it falls as rain. It then flows as a stream towards a valley turning the water-wheel of a mill on its way. Describe the energy transformations which take place.

5. A screw-jack which has a velocity ratio of 40 requires an effort of 100 lb. weight to raise a load of 1,200 lb. weight. What is the efficiency of the screw-jack?

6. What do you understand by the term friction? Distinguish between sliding friction and limiting friction.

The wheels of a locomotive sometimes skid on the line when the engine begins to pull a heavy train from rest up a slight gradient, but once the train is moving the skidding stops. Explain this and state what the driver does when the wheels skid.

7. A 10-stone man carrying a pack of 28 lb. climbs up a mountain 1,100 ft. high in 2 hours. Calculate the average horse-power exerted by the man.

8. A 12-stone man sitting 1 ft. from the end of a 12-ft. see-saw balances a boy of 8 stones sitting at the other end. How far is the man from the point of balance?

9. A uniform rod, AB, is free to rotate in a vertical plane about B. A 10-lb. weight hangs at A. What vertical force must be applied at the mid-point of the rod to keep it stationary? (Neglect the weight of the rod.)

10. Draw a diagram of a pulley having a velocity ratio of 4. If this pulley raises a 50-lb. weight with an effort of 20 lb. weight what is the mechanical advantage and the efficiency?

FORCES ACTING ON A BODY
CENTRE OF GRAVITY
RESULTANT FORCE

If you hold a pole, with one end in your hand and the other touching a wall, and push at your end in the direction of the pole, the force you apply will be transmitted through the pole and take effect on the wall. In general there are three things of which we must take account in considering forces:

1. The point of application of the force, that is, exactly where the force acts.
2. The direction of the force, including its line of action.
3. The magnitude of the force, that is, whether it is a force of 2 lb. weight or 3 lb. weight.

Suppose a friend came to your end of the pole and pushed in the same direction at the same time as you. The force exerted on the wall would be one force which was the sum of the two forces applied. If you pushed with a force of 25 lb. weight and your friend with a force of 10 lb. weight, the force acting on the wall would be one of 35 lb. weight. This force is called the *resultant* of the two forces, each of which is called a *component* of the total force.

Now suppose that while you were pushing with a force of 25 lb. weight your friend pulled in exactly the opposite direction with a force of 10 lb. weight. The resultant force on the wall would be the difference of the two components, namely, 15 lb. weight.

These are forces acting in the same straight line. What happens when the lines of action are different?

PARALLELOGRAM OF FORCES

When two boys kick a football at the same instant but in different directions, the football travels in neither of these directions, but in one quite different. The force which causes

the ball to move is the resultant of the two forces (kicks) applied by the boys. It is in a different direction; is it also different in magnitude? This can easily be discovered by experiment.

Expt. 232. A drawing-board on which is pinned a sheet of paper is supported vertically (Fig. 174). P_1 and P_2 are light pulleys round which pass thin cords attached to a small ring O. Another cord is attached to O and
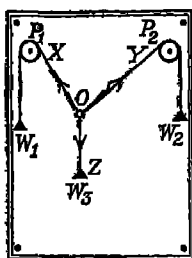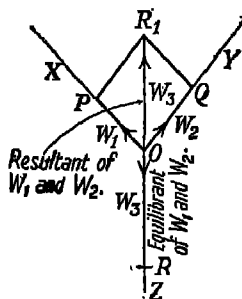
FIG. 174.

hangs vertically supporting a weight, W_3. W_1 and W_2 are weights attached to the cords from the pulleys. The ring O, which if small may be considered a point, is at rest. The three forces acting on it must balance. Consider the two forces W_1 and W_2. One acts in the direction OX and the other in the direction OY. As O is at rest, the force W_3 acting in the direction OZ must balance them. This force is called the *equilibrant* of W_1 and W_2.

Mark on the paper the position of the point O. Draw lines from O in the direction of OX, OY, and OZ. Note the weights W_1, W_2, and W_3, and remove the cords so that you can draw properly on the paper. Along OX mark a length OP to represent the force W_1 to scale. Using the same scale mark OQ along OY and OR along OZ to represent the forces W_2 and W_3 respectively. Produce ZO to R_1 making $OR_1 = OR$. From P and Q draw lines parallel to OY and OX respectively. You will find that these lines intersect at, or very nearly at, the point R_1.

FIG. 175.

OR_1 represents in magnitude and direction the resultant of the forces W_1 and W_2. If the force W_3 is removed then there are two forces W_1 and W_2 acting on O in the directions OX and OY. These forces act as if there were one force acting on O, of magnitude W_3, and in the direction OR_1. OPR_1Y is a parallelogram with OR_1 the diagonal. If two forces acting at a point O are represented in magnitude and direction by the two sides of the parallelogram OP and OQ, their resultant is represented in

magnitude and direction by the diagonal of the parallelogram OR_1. This theorem is known as the *parallelogram of forces*.

RESOLUTION OF A FORCE

We have seen that two forces acting on a body can be considered to have one resultant force, and this can be found by applying the principle of the parallelogram of forces. Consider a truck on a railway line being pulled by a horse walking
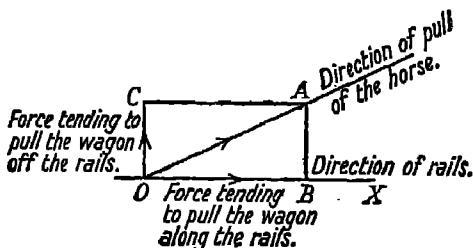
FIG. 176.

outside the track. Clearly the line of action of the force is not in the direction of the railway line which is the only direction in which the truck can move. The single force acting on the truck can be resolved into two forces, one tending to move the truck along the rails and the other tending to pull it off the rails, that is, in a direction at right angles to the rails. The useful portion of the force is the one which tends to move the truck along the rails. Suppose the horse is pulling with a force of 100 lb. weight at an angle of 25° to the rails. By using the principle of the parallelogram of forces the useful portion of this force can be found. In Fig. 176 OX represents the direction of the rail, OA represents in magnitude and direction the pull of the horse. AB is perpendicular to OX. OB represents in magnitude and direction the component acting along the rail. It can be found by calculation or graphically. OC represents in magnitude and direction the force tending to pull the truck off the line. This is resisted by the reaction of the rails.

In many cases the force applied to a body is not always in

the direction required. There are practical difficulties which prevent this.. A horse walking on the tow-path pulling a barge tends to pull it to the side of the canal as well as along the canal. A barrow is pushed in the direction of the handles, but only a part of this force pushes it along the ground. There is also a vertical component tending to force the barrow into the earth. The resolution of a force is of great practical importance.

Fig. 177. Resolution of a force.

Fig. 178. Resolution of a force.

CENTRE OF GRAVITY

. **Expt. 233.** Balance a ruler on the end of your finger, or, if you have enough patience, on the point of a pair of compasses.

Consider the forces acting on it when it is balanced. There is the force of gravity acting on every particle of the ruler, but all these small forces result in one force, the resultant, which acts vertically downwards through the point at which the ruler is balanced. The finger also exerts an upward vertical force through the same point and this must be equal to the downward force, otherwise the ruler would move up or down. The point through which the line of action of the weight of a body always passes, no matter what the position of the body, is called the *centre of gravity* of the body. It is sometimes described as the point at which the whole weight of the body may be supposed to act.

Expt. 234. Push a book or other similar object (a piece of cardboard or a thin piece of wood) over the edge of a table. Proceed very slowly and stop exactly when the book is ready to overbalance and fall to the floor. Draw a pencil line on the book showing the position of the edge of the table. Push the book again from another direction and draw another line marking the edge of the table. There will now be two intersecting lines on the book. The point of intersection is the centre of gravity of the book. Try to balance the book on the finger at the intersection of the lines. If you push the book from still another direction you will obtain another line which will go through the same point, or very nearly through it, according to the care you take.

Strictly speaking the point you found is not the centre of gravity, because the book being a solid body has the centre of gravity actually inside it. The point you found is vertically under the centre of gravity when the book is lying on a horizontal surface. If instead of using a book you use a piece of cardboard, the point you find is very near the actual

Fig. 179. Centre of gravity of a piece of cardboard.

centre of gravity, because the cardboard is very thin.

To find the Centre of Gravity of Various Bodies

Expt. 235. To find the centre of gravity of a rectangular piece of cardboard. Make a small hole anywhere near the edge of the cardboard and suspend it by means of a smooth pin or nail through the hole. Use a plumb-line in the way shown in Fig. 179 to draw on the cardboard a vertical line through the point of suspension. Now suspend the cardboard from another position near the edge and draw another line. The centre of gravity is at the intersection of the two lines. Test by trying to balance the cardboard on a pencil point. If you suspend the cardboard from many points all the lines should pass through the same point. They will probably not do this exactly, but by careful work they may be made to do so very nearly.

If a body is at rest and is suspended from one point, then the centre of gravity of the body and the point of suspension are in the same vertical line, provided that the body is not acted on by any force other than gravity. This gives a practical method of finding the centre of gravity of a body as in the above

experiment. It is also true that if a body is supported in any way at a single point and is at rest, then the point of support and the centre of gravity are in the same vertical line.

Draw the diagonals of the rectangle and note that the intersection of the diagonals is the same point you found for the centre of gravity. The centre of gravity of any parallelogram is the intersection of its diagonals.

This practical method may be used to find the centre of gravity of any thin, flat body, no matter what the shape. Use it to find the centre of gravity of a triangle, and of an irregular-shaped lamina. The centre of gravity of a triangle is at the intersection of the medians.

The centre of gravity of a brick is at the point of intersection of the diagonals of the brick (not the diagonals of one of its faces). Obtain a light wooden brick and try to balance it on one finger. Note that every time you get it balanced (or nearly balanced) your finger is vertically below the centre of gravity.
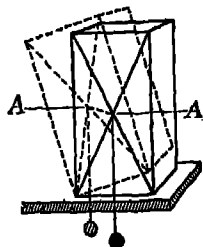
FIG. 180. Stability of a block.

The centre of gravity of a sphere is its centre, and of a hollow cone, a point on the axis one-third of the way up from the base.

STABILITY

No doubt you have noticed that some objects are much more easily upset than others. Even people who have not studied mechanics know when an object looks unstable. Perhaps they could not put into words why a particular object is unstable, but they have a good idea of the cause of instability. Often by looking at an object we decide that it is unstable. We say that it is 'top heavy'. But appearance is not good enough. A body may look unstable and yet be perfectly stable. The Leaning Tower of Pisa, for example, looks unstable, but it has stood for hundreds of years. Let us try to understand something about this important question of stability.

Expt. 236. Obtain a block of wood about the same size and shape as an ordinary brick. Stand it on a flat piece of wood on a table with the greatest length of the block vertical (Fig. 180). Draw the diagonals of one face. The point of intersection of the diagonals (*A*) is not the centre of gravity of the block but it is the same height as the centre of gravity. For the purpose of this experiment we may consider it to be the centre of gravity. Drive a small nail into the block at this point, and fasten a plumb-line to the nail. Raise one side of the wood on which the block stands so that the block is tilted. Note that the plumb-line travels nearer to the edge of the block. Gradually tilt the
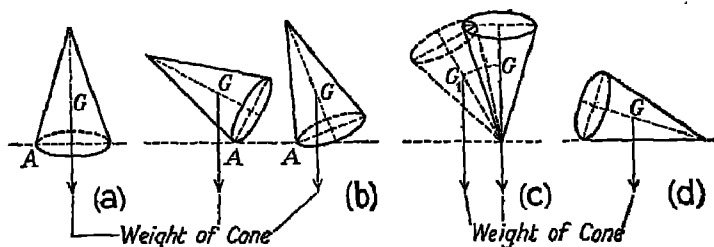
FIG. 181. Stability of a cone.

block until it overturns. What is the position of the plumb-line when this occurs?

The block overturns when the vertical from the centre of gravity falls outside the base.

Consider a hollow cone. The centre of gravity is in the position shown in Fig. 181(*a*). When standing on its base it is stable, or, as we say, in *stable equilibrium*. Any tilting of the cone causes its centre of gravity to rise, and if left alone the cone will fall back on its base. Consider the forces acting on the cone in this position. The cone is free to rotate about the point *A*. A force, the weight of the cone, acts vertically downwards at G. If no other force is acting on the cone it must fall back on its base. Now suppose the cone is tilted so that the perpendicular from its centre of gravity falls outside the base (Fig. 181*b*). It is free to rotate about *A* but its weight acting at *G* this time causes the cone to fall over. The centre of gravity thus tends to become as low as possible. When it is in its lowest position the body is in stable equilibrium.

If you could stand a cone on its apex (Fig. 181*c*) any tilting

would cause the centre of gravity to fall and there is no force acting to bring it back to its original position. Indeed, there is a force, its weight, acting at the centre of gravity, tending to make the cone rotate about its apex and so fall on its side. Looking at it from the other point of view, the base is very small (a point, the apex). The least tilting of the cone causes the vertical line from the centre of gravity to fall outside the base. In this position the cone is said to be in a position of *unstable equilibrium.*

If the cone is lying on its side it may be rolled without altering the height of its centre of gravity. In this position it is said to be in *neutral equilibrium.* A uniform sphere is always in neutral equilibrium.

Fig. 182. A farmer's two-wheeled cart.

EXAMPLES ILLUSTRATING STABILITY

A farmer's two-wheeled cart, loaded in the ordinary way, may be stable, but the way some of them are loaded high with hay makes them not very stable because the centre of gravity is raised, and the base (the line joining the points of contact of the wheels with the ground) is not very wide (Fig. 182).
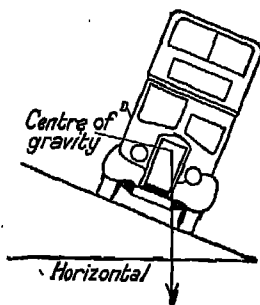
Fig. 183. Stability of a bus.

If the cart, so loaded, is tilted by one wheel passing over a large stone, there is danger that the vertical line from the centre of gravity may fall outside the base.

The same applies to a bus (Fig. 183). The centre of gravity is low because the heavy parts like the engine and tanks are kept as low as possible. Unless the centre of gravity is very low it is dangerous to load the top deck with passengers when there are no passengers on the lower deck. That is why standing is not allowed on the upper deck. The stability of buses is

subject to government regulations and control, and if they are not stable under the specified conditions they are not allowed on the roads.

The stability of ships is subject to the inspection of the Board of Trade. When a ship is not carrying cargo it must carry ballast. This keeps the centre of gravity low. Some ballast has to be carried, cargo or no cargo, by most ships. This is usually carried near to the keel.

People hiring rowing boats, especially light skiffs, are warned not to stand up in them. This is because a person standing up raises the centre of gravity and a slight tilt might cause the boat to capsize. Wise people, if they must change seats in a light boat, do so by crawling to their new places, as this keeps the centre of gravity low.

Have you noticed that when you are carrying a heavy weight in one hand there is a tendency for you to stretch the unused arm as far as possible on the other side of the weight? This is an attempt to 'restore the balance' or to bring the centre of gravity back to the centre line of the body (Fig. 184).

Fig. 184.

A tight-rope walker must keep his centre of gravity as nearly as possible vertically above the wire on which he is walking. He sometimes carries a pole which enables him to do this. By changing the position of the pole he can better control the centre of gravity of the pole and himself.

Some people have the bad habit of trying to balance themselves on a chair tilted backwards. It is a difficult feat as the vertical line from the centre of gravity easily falls outside the base. It can be done by altering the position of one's legs a

little so as to bring the centre of gravity first to one side and then to the other of the line representing the base.

We have said that some things look 'top heavy' or unstable when in fact they are not so. If a hemisphere of lead is

Centre of Gravity
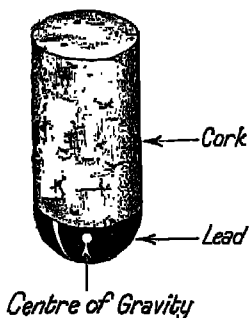
Fig. 185.

fastened to a cylinder of cork of the same diameter (Fig. 185) and then painted so that the whole appears to be made of the same material, it will rest with the lead on the table and it will be very difficult to knock it over. It will, however, look unstable, because all the material appears to be the same, and it will be thought that the centre of gravity is high, whereas in fact it is low, the lead being much denser than the cork. This device is employed in making many kinds of amusing toys.

Try to balance a walking-stick vertically on your finger. Some people can do it on the nose. You cannot balance it unless the finger is moved at least slightly. The base of the walking-stick is very small and the vertical from the centre of gravity very soon falls outside the base. This may be corrected by moving the base so that the vertical falls within it. Watch your movements as you balance the stick, and try to find out exactly what you are doing.

Have you ever tried to walk on stilts? If not, try it, for it is great fun. You will find that you must keep the stilts moving or you will fall. Try to think out why.

BALANCE IN ANIMALS

How does an animal manage to keep the 'right side up'? Some animals, including man, look unstable mechanically. We are so used to seeing men stand on their feet without difficulty that perhaps we do not realize that mechanically man is not a very stable body. A man is much more stable when he is standing with his feet astride, but even then he would not be able to 'keep his balance' if he had not the control of his limbs. An unconscious man would fall when placed on his feet. The

fact that a man can stand on his feet depends on more than mechanical principles. The body, besides being a machine, possesses 'life'. If a cat is held in the air by its four feet with its back downwards and then released, it lands on its feet. It makes a remarkable turn in the air. A dead cat treated in this manner does not land on its feet. How does a fish keep the 'right side up' even in the dark? When it is dead it usually floats ventral side upwards. We can understand better the mechanism in animals by studying the simpler animals first.

Aperture
of statocyst

Statocyst
&
sensory hairs
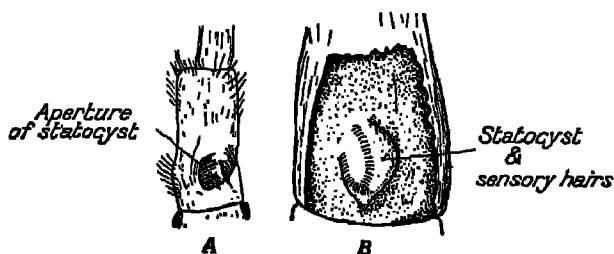
FIG. 186. Statocyst from a crayfish.
A. External view; B. wall cut away.

THE STATOCYST

Even jellyfish have a mechanism to enable them to keep the right side up in water. The mechanism, however, is best studied in such animals as the shrimp, prawn, crayfish, or lobster. Near the base of each of the smaller pair of feelers of the shrimp is a small chamber called a statocyst (Fig. 186). From the lining of the chamber project numerous very stiff hair-like processes. In the statocyst are some grains of sand which rest on the hairs, but as the latter are fixed and the sand grains are free to move, the grains are in contact with different hairs according to the position of the shrimp. The hairs are connected with nerves which control the various muscles causing movement (Ch. XLV). If a shrimp swimming normally turns over slightly, the grains of sand (called statoliths) move and so come in contact with other hairs. The nerves thus

stimulated pass the message to the muscles controlling the
movements of the shrimp, which thus regains its normal
position. Fig. 187 illustrates diagrammatically how a statocyst
works. It is not a picture of what one looks like. A shrimp does
not correct its position consciously. Even the higher animals,
ourselves included, do not do that. The correction is automatic.
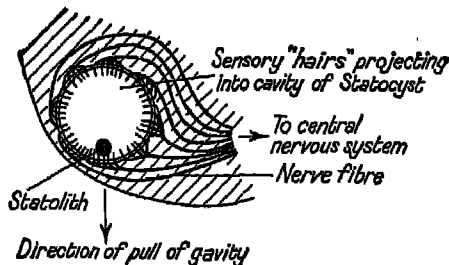
FIG. 187. Diagrammatic representation of how
a statocyst works.

A MODEL OF A STATOCYST

A working model of a statocyst may be very simply made. Cut an
india-rubber ball and insert some small copper shot. Put metal pins
through various areas of the ball and connect all the pins in one small
area to a bulb and battery. Similarly connect up each of the other
side areas to separate bulbs. When the ball is moved the copper shot
roll about inside the ball and fall on the metal pins, thus completing
a circuit and causing one or other of the bulbs to light. This is not a
complete model, as it represents only one part of the operation. There
is nothing to represent the mechanism of correcting the position of
the ball. The message goes to the nerves but these send no message
to the muscles to correct the balance.

BALANCE IN SHRIMPS

· It is possible to replace the sand in the statocysts of shrimps
by iron filings. When this is done the shrimps act normally
as regards balance. When, however, a magnet is placed above
them, the filings are attracted in the direction of the magnet
and are not controlled by gravity alone as before. (Gravity
still acts on them but the magnet is the stronger force.) The

filings, acting as statoliths, come in contact with hairs different from those they would have touched had gravity been acting alone. Shrimps may be made to swim upside-down by controlling the iron filings with a magnet.

BALANCE IN HIGHER ANIMALS

In fish, frog, and man, the organ of balance is in the inner ear. It follows the same principles described above, but is more complicated. The statoliths consist of calcium carbonate secreted by the animal. We shall leave the description of this organ until the function of the ear is dealt with in Sound (Ch. XXXV). Sight is also of importance in maintaining balance, but it can be dispensed with if the balancing mechanism of the ear is in order. A blind man, otherwise sound, has no difficulty in maintaining his balance. A man is able to maintain his balance by the delicate control of many muscles. Sight and the balancing mechanism of the ear help him to keep nervous control of all these muscles. It is all done subconsciously. A man in good health is not conscious of any effort to maintain his balance. It is the same with learning to ride a bicycle. After we have once learned we are not conscious of any effort at control.

Not all animals have statocysts. An earthworm has no such apparatus, and animals like sea anemones which are attached to rocks do not need a balancing apparatus.

QUESTIONS

1. State carefully what you understand by the term the centre of gravity of a body. How does the position of the centre of gravity of a body affect its stability? Illustrate your answer by reference to a double-deck bus.

2. Two forces, 10 lb. weight and 15 lb. weight respectively, act at a point. The angle between their lines of action is 60°. Find, by means of a geometrical construction, the magnitude and direction of the resultant.

3. A force of 100 lb. weight is resolved into two components, one of 80 lb. weight acting at 30° to the direction of the force. Find, by means of a geometrical construction, the magnitude and direction of the other force.

C C

4. A horse on a tow-path pulls a barge along a straight canal. The tow-rope is at an angle of 30° to the canal and the horse exerts a force of 100 lb. weight. What force tends to pull the barge (*a*) along the canal, (*b*) towards the bank?

XXII

DENSITY AND SPECIFIC GRAVITY

DENSITY

WE say sometimes that lead is heavier than cork. What exactly do we mean? It is obvious that a small piece of lead is not heavier than a large piece of cork. We mean that if we have a piece of lead and a piece of cork of the same size then the lead is the heavier. So instead of saying that lead is heavier than cork we ought to use a word which means 'heavier for the same size'. This word is 'denser' and we say that lead is denser than cork.

Instead of thinking of the weights of equal volumes of lead and cork we could equally well consider the volumes of equal weights of those substances. The volume of a pound of lead is much smaller than that of a pound of cork. This method of comparison is not so convenient as the other.

The *density* of a substance is defined as the mass of unit volume. We find the mass by weighing, and in practice we always think of the density as weight per unit volume. It does not matter what the unit of volume is. Sometimes one unit is more convenient than another for a particular purpose. Thus the density of water may be expressed as:

(1) 1,000 oz. per cu. ft.
(2) 10 lb. per gallon.
(3) 1 gm. per c.c.

When the metric system was introduced into France the gram was defined as the mass of 1 c.c. of pure water at 4° C.

We may write down the relation between density, mass, and volume in the form

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}.$$

If any two of these quantities are known the third can be found.

Expt. 237. **To find the densities of some solids.** The densities of some solids can be found very easily. All that is required is the weight of a known volume. Select a suitable solid and weigh a known volume on a chemical or spring balance. If the solid has a regular shape you can measure it and calculate its volume by arithmetic. The density is found by dividing the weight by the volume. If the solid is irregular in shape, find its volume by immersing it in water in a graduated cylinder. The rise in the level of the water gives you the volume. Remember that the wider the cylinder the less accurate your reading may be. If the solid is made of wood or other substance which will not sink, make it do so by holding it under water with a thin wire, the volume of which is so small that it can be neglected. Another way is to sink it under water by a piece of iron, the volume of which has been found either experimentally or by calculation.

Expt. 238. **To find the density of a liquid.** Weigh a dry, clean beaker. Into this run from a pipette or burette a known volume of the liquid. Weigh again. From these facts calculate the density.

Try to discover for yourself ways of finding the densities of substances.

RELATIVE DENSITY. SPECIFIC GRAVITY

It is often convenient to compare the densities of different substances. An easy way of doing this is to compare the densities of all the substances with that of a convenient standard substance. For solids and liquids this substance is water; for gases it is usually hydrogen, the 'lightest' gas known, although air, a commoner gas, is often used. Suppose the density of glass is 2·6 gm. per c.c. Clearly the density of glass is 2·6 times that of water, which is 1 gm. per c.c. This is more briefly expressed by saying that the Relative Density of glass is 2·6. Another name for Relative Density is *Specific Gravity* (usually contracted to S.G.). If we had taken the density of glass in oz. per cu. in. instead of in gm. per c.c., and the density of water also in oz. per cu. in., we should have obtained the same number, 2·6, for the relative density. The convenience of the metric system is that in it the same number is used in stating the density and the relative density; the density of glass is 2·6 gm. per c.c., the relative density is 2·6.

$$\text{S.G. of a substance} = \frac{\text{Density of the substance}}{\text{Density of water}},$$

$$= \frac{\text{Wt. per unit vol. of the substance}}{\text{Wt. per unit vol. of water}},$$

$$= \frac{\text{Wt. of any vol. of the substance}}{\text{Wt. of an equal vol. of water}}.$$

The S.G. of a substance may be found by first finding the density as in Experiment 237, but there are more convenient methods.

To Find the Specific Gravity of a Substance by Means of the Specific Gravity Bottle

Expt. 239. To find the S.G. of a liquid (turpentine). The S.G. bottle (Fig. 188), sometimes called a Density Bottle, is a small flask, of about 50 c.c. capacity. It is fitted with a glass stopper through which runs a narrow hole. Weigh the bottle empty. Fill the bottle with turpentine and then insert the stopper. Some turpentine will come out through the hole. Make sure that there are no air bubbles under the stopper and that the hole is full of turpentine. Wipe the outside of the bottle dry. Hold the bottle by the neck so that the heat of the hand will not warm the whole bottle. Weigh the bottle and its contents. Empty the bottle and wash it out well with water. Fill with water, taking the same precautions as before. Weigh.
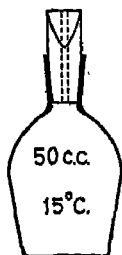
Fig. 188. A density bottle.

50 c.c.
15°C.

		Grams
Weight of empty bottle .	. . . . .	a
Weight of bottle + turpentine .	. . . .	b
Weight of bottle + water	. . . . .	c
Weight of turpentine .	. . . .	$b-a$
Weight of an equal volume of water	. .	$c-a$

$$\text{S.G. of turpentine} = \frac{(b-a)}{(c-a)}.$$

In the same way the specific gravities of other liquids such as salt solution, paraffin, methylated spirit, may be found.

Expt. 240. To find the S.G. of a solid insoluble in water. The solid must be small enough to go into the S.G. bottle or it must be broken up into pieces small enough to go into the bottle. (1) Weigh the solid. (2) Weigh the solid and the bottle full of water on the same

pan. (3) Place the solid in the bottle and fill with water. Take care that there are no air bubbles. Weigh.

Grams

Weight of solid   .   .   .   . a

Weight of solid+bottle full of water   .   .   . b

Weight with solid in the bottle full of water   .   . c

Weight of water displaced by the solid   .   .   . $b-c$

$$\text{S.G. of the solid} = \frac{a}{(b-c)}.$$

Other methods of finding the S.G. of substances will be dealt with after the study of Archimedes' Principle.

ARCHIMEDES' PRINCIPLE

When objects are thrown into water some sink and some float. Those which float are said to be 'lighter' than those

which sink. It is clear that gravity must be acting on a floating body as on all other bodies, attracting it towards the centre of the earth. As it does not move, there must be an equal thrust acting vertically upwards. This thrust must be exerted by the water.

Does this force also act on bodies which sink? If so, why do some sink and some float? Our experience suggests that this force does act on bodies which sink, that is, on those which are totally immersed. This fact can be tested experimentally.

Expt. 241. Is there an upward thrust when a body is totally immersed? Arrange the apparatus as shown in Fig. 189. Attach an object to the spring balance, which is supported by a retort stand. Note the weight of the object in air. Now allow the object, still attached to the spring balance, to sink in some water in a beaker. The beaker may be supported by blocks of wood of the requisite thickness. Note the weight of the object in water. Repeat the experiment with other objects which sink in water.

FIG. 189. Weighing an object immersed in water.

We now know that when a body is partially or totally immersed in water there are at least two forces acting on it, (i) the force due to gravity, and (ii) a force due to water which

acts in the opposite direction to that of gravity. Let us now find the magnitude of this upward thrust due to the water.

Expt. 242. To find the upward vertical thrust when a body is totally immersed in water (Bucket and Cylinder experiment). Fig. 190. A metal cylinder just fits into a cylindrical bucket. The cylinder is suspended from the bucket which is itself suspended from the balance, as shown in the figure. Both are weighed in air. A beaker of water standing on a bridge is so placed that the cylinder is immersed in the water. The two sides no longer balance. Water is poured into the bucket until it is full. This restores the balance.

The upthrust is equal to the weight of water placed in the bucket. The volume of the water placed in the bucket is of course equal to the volume of the cylinder. The upward thrust is equal to the weight of the water displaced.
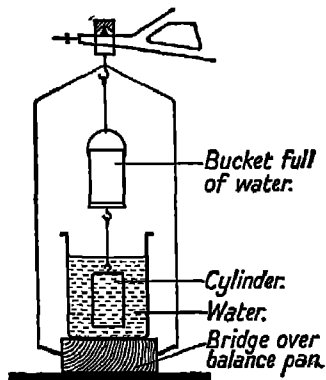
FIG. 190. The bucket and cylinder experiment.

This fact is known as the *principle of Archimedes.* It may be stated thus: When a body is immersed in a liquid, there is an upward thrust which is equal to the weight of the liquid displaced.

ARCHIMEDES was a Greek. He is considered one of the greatest philosophers of all time. He was a mathematician, but discovered many new principles in the branch of science we now call Physics. He was also noted for his inventive genius.

We can now understand why some bodies float, and some sink, when put in water. If the weight of the water displaced by a completely immersed body is less than the weight of the body, in other words, if the S.G. of the body is greater than 1, the body sinks because the upthrust is less than the weight. On the other hand, if the weight of the water displaced is greater than the weight of the body (i.e. S.G. less than 1) the upward thrust is greater than the force of gravity acting on the body. The latter therefore rises until it comes to the surface. Part of

the body is then out of the water, so that the upward thrust is not so great, because the upward thrust is always equal to the weight of the water displaced. Finally a stage is reached when the upward thrust is equal to the weight, and the body remains at rest, floating on the water.

Expt. 243. To illustrate the fact that the principle of Archimedes applies also to floating bodies. Take a graduated cylinder containing water. Note the level of the water. Float in it a test-tube containing some lead shot to keep the tube floating vertically. Without the shot the upward thrust of the water would be so great that the test-tube would not sink deeply enough. Note the new level of the water. The volume of the water displaced by the floating test-tube and shot is given by the difference between the two readings. The weight of the water displaced is thus known. Remove the test-tube and, after drying it, weigh it and the shot. This weight will be equal to the weight of the water displaced.

The Principle of Archimedes is quite general and applies to gases as well as liquids. It may be stated in the most general terms as follows:

When a body is totally or partially immersed in a fluid (i.e. a liquid or gas) there is an upward thrust equal to the weight of the fluid displaced.

TO FIND THE SPECIFIC GRAVITY OF SOME SUBSTANCES,
USING THE PRINCIPLE OF ARCHIMEDES

Expt. 244. To find the S.G. of a solid which sinks in water. Weigh the solid in air, then weigh it suspended in water. The arrangement used in Fig. 189 may be used.

							Grams
Weight in air	.	.	.	.	.	.	x
Weight in water	.	.	.	.	.	.	y
Weight of water displaced	.	.	.	.	.	$x-y$

$$\text{S.G.} = \frac{\text{Wt. of any vol. of the solid}}{\text{Wt. of an equal vol. of water}} = \frac{x}{x-y}.$$

Expt. 245. To find the S.G. of a solid which does not sink in water (a piece of paraffin wax). Take a piece of metal heavy enough to sink the wax in water. Weigh the metal in water. Fasten the metal to the wax by means of thread and arrange it as shown in Fig. 191, so that the metal is in water and the wax in air. Weigh in this position. Now weigh both the metal and wax together in water.

		Grams
Weight of metal in water	. . . . .	x
Weight of metal in water+wax in air	. . . .	y
Weight of metal in water+wax in water .	. . .	z
Weight of wax	. . . . .	$y-x$
Weight of water displaced	. . . . .	$y-z$

$$\text{S.G.} = \frac{y-x}{y-z}.$$

No attempt should be made to remember this result but the method should be studied.

Expt. 246. To find the S.G. of a liquid (a salt solution). Weigh a piece of glass or metal in air, water, and salt solution.

	Grams
Weight of glass in air .	a
Weight of glass in water .	b
Weight of glass in salt solution	c
Weight of water displaced by the glass . . .	$a-b$
Weight of salt solution displaced by the glass . .	$a-c$

These are weights of equal volumes.

Therefore S.G. of salt solution $= \dfrac{a-c}{a-b}.$

Fig. 191. To find the specific gravity of a solid 'lighter' than water.

Again remember the method and not the result.

HYDROMETERS

It is often important to know the S.G. of a liquid. The S.G. of beer and spirits is an indication of the amount of alcohol in them, and the Excise Authority tests them for this by finding the S.G. The S.G. of solutions of ammonia and sulphuric acid is an indication of their strength. We can tell when an accumulator needs charging by finding the S.G. of the acid in it. Oils used in industry are classified as heavy or light according to their specific gravities. The S.G. of skimmed milk is greater than that of unskimmed milk, because the cream which rises to the surface is less dense than the remainder of the liquid, and if removed there is a rise in the S.G. of the whole liquid. If,

however, water is added after the milk has been skimmed the S.G. is again lowered. The removal of the cream can thus be disguised. Fortunately it cannot be disguised from the chemist.

It is thus convenient to have an instrument which will give the S.G. of a liquid very quickly. Such instruments are called
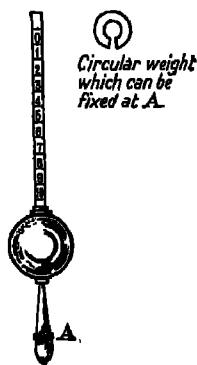
FIG. 192. Sykes's hydrometer.
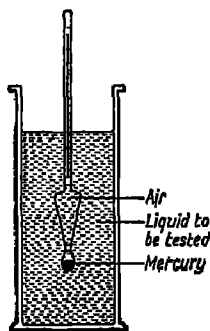
FIG. 193. Twaddell's hydrometer.
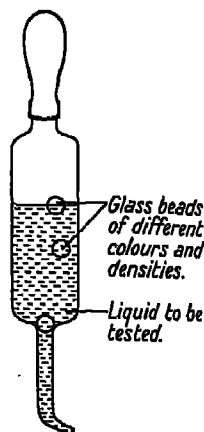
FIG. 194. A cheap and convenient hydrometer.

hydrometers. They are designed on the principle of the floating test-tube loaded with shot, which we used in Experiment 243. A hydrometer consists of a floating body weighted so that it floats upright. The upper part, the stem, is thin and is graduated so that the S.G. of the liquid in which it floats is given by the reading on the scale at the liquid level. As the specific gravities of liquids vary widely, hydrometers may be differently weighted to suit particular liquids. That used for milk is called a *lactometer.* It would be useless for finding the specific gravities of heavy oils as it is graduated for milk only, the S.G. of which varies between 1·015 and 1·033. That used by the Excise Authority is called the Sykes's hydrometer (Fig. 192). The Twaddell hydrometer (Fig. 193) is also a common type.

Others are very much simplified, especially when they are used to find if a given liquid has a particular S.G. or not. One consists of a glass tube like a fountain-pen filler, containing a number of coloured balls of varying densities which float or sink according to the S.G. of the liquid drawn into the tube. This type is commonly used for testing the acid of an accumulator (Fig. 194). This principle may be illustrated by the following experiment.

Expt. 247. Place an egg in a tumbler of water. The egg sinks. Add strong salt solution to the water in the tumbler and stir. The egg floats. Try again, this time to add just enough salt solution so that the egg neither sinks nor floats. It is possible but not easy.

SHIPS

Ships float in accordance with the Principle of Archimedes. The upward thrust due to the water is equal to the weight of the water displaced by that part of the ship under water. Where the surface of the water touches the side of the ship is called the water-line. When a ship is loaded with cargo it sinks more deeply into the
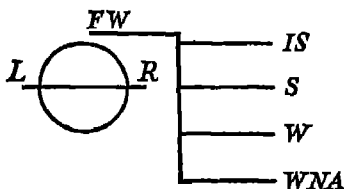
Fig. 195. The Plimsoll line.

water; in other words, the water-line rises. It is clear that if the water-line rises too high there may be danger for the ship in a rough sea. For this reason, the Board of Trade in Great Britain, and some similar authority in other countries, insist on a water-line below which the ship must not be loaded. Fig. 195 illustrates the markings made on the side of a ship. The line LR passing through the circle is called the *load* line or the *Plimsoll line*. It indicates the depth to which the ship may be immersed by loading. The letters L.R. mean Lloyd's Register of Shipping. A higher line FW is allowed when the ship is sailing in fresh water. A ship loaded to this line in fresh water rises to the line LR as it passes out into the sea, owing to the greater density, and therefore the greater upthrust, of salt water. The other lines are: IS, Indian

Summer; *S,* a continuation of the Plimsoll line; *W,* Winter, and *WNA,* winter in the North Atlantic.

The Plimsoll line was made compulsory in 1876. Before then ships were often overloaded and there were many losses. Samuel Plimsoll successfully agitated until a safe load line was made compulsory.
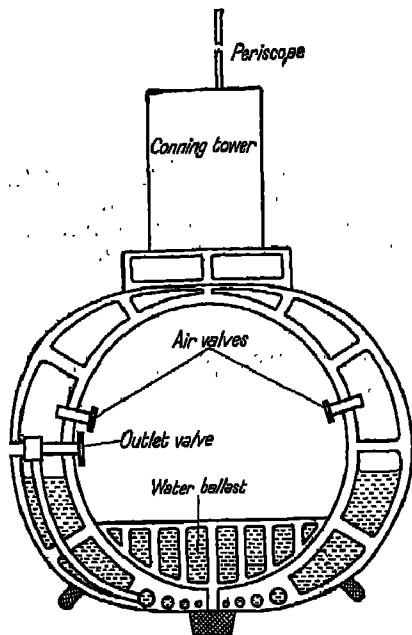
FIG. 196. Diagram of the section of a submarine.

SUBMARINES

Submarines normally float on the surface like other ships. They are, however, fitted with tanks which are full of air when the ship is on the surface but which may be filled with water by opening valves when necessary. The density of the ship is thus increased and it sinks below the surface. The depth is regulated in two ways: (i) by increasing or decreasing the amount of water in the tanks, and (ii) by the operation of a horizontal rudder, so that with a driving screw the ship is steered upwards or downwards. When it is desired to rise to the surface, pumps worked by compressed air force the water out of the tanks (Fig. 196).

The tanks of a submarine may be compared to the swim-bladder of a fish. This is filled with air or other gas secreted from the blood and so makes the weight of water displaced by the fish equal to the weight of the fish. Little or no effort is required for the fish to rise or sink.

BALLOONS

The Principle of Archimedes applies to gases as well as to liquids. Any body in the air is subject to an upward thrust which is equal to the weight of the air displaced by the body. In the case of a solid the upthrust is so small compared with the weight of the solid that it can generally be neglected. This we did in our specific gravity experiments using the Principle of Archimedes. We assumed the weight of a body in air to be the same as its weight in a vacuum. This is not the case, but the difference in the case of solids and liquids is very small.

In the case of a balloon it is the upward thrust of the air which causes the balloon to rise. A balloon is essentially a bag containing a gas much less dense than air. This gas is usually hydrogen, the lightest (least dense) gas known. Sometimes helium is used, since although it is denser than hydrogen it has the advantage of being non-inflammable.

Consider the vertical forces acting on a balloon. There is the force due to gravity which acts on the hydrogen and on the envelope. The former, however, is very light and the envelope is made as light as possible. The balloon is large and displaces a great volume of air. The upward thrust is greater than the downward force of gravity, and the balloon therefore rises. The air, however, is less dense as the height above the surface of the earth increases (p. 42), and the balloon in time comes to a region where the weight of the balloon is exactly equalled by the weight of the air displaced. It then stops rising. If it is desired to descend, some hydrogen is allowed to escape from the envelope, thus decreasing the weight of the air displaced and causing the balloon to descend. The modern 'lighter than air' airship carries a heavy cargo. It is therefore essential that the upthrust of the air should be very great. This can only be done if a very large volume of gas much less dense than air displaces an equal volume of air.

QUESTIONS

1. Distinguish between (a) the density, and (b) the relative density of a substance. Describe a simple way in which you could find the density of an irregularly shaped piece of paraffin wax.

2. State the Principle of Archimedes.

A body weighs 5 gm. in air and 4 gm. in water. What is the density of the body?

The weight of the body in another liquid was 3·8 gm. What is the relative density of the liquid?

3. A rectangular prism of wood, 9 cm. long, 4 cm. broad, and 3 cm. deep, floats on water, 1 cm. of the block being above the level of the water. A 20 gm. weight is placed on the block. How much farther does it sink into the water?

4. What is a hydrometer and for what is it used? Describe any common type of hydrometer and explain its use.

5. What is the purpose of the Plimsoll line on a ship? Why is it different for fresh water and sea water?

6. Describe how a submarine submerges and surfaces. Explain carefully the principle involved.

7. Explain why a hydrogen balloon descends when gas is allowed to escape from the envelope. Will it continue to descend to the surface of the earth if no more gas is allowed to escape? Give a reason for your answer.

XXIII

FLUID PRESSURE

PRESSURE OF WATER

THERE is a common saying: 'Water finds its own level', as though water were a living thing and always made an effort to make its surface level. We know that this is not so, but we also know from our experience that the surface of water is usually level and when something happens to make it otherwise the water flows until it is level again. How then does water flow uphill to a house situated there? If it can do that why does it not flow higher to the top of the house when there is a fire there and so save the fireman the necessity of pumping it? These questions cannot be answered properly until we have studied the pressure of water. Let us first see what we mean by the term pressure.

PRESSURE AND TOTAL THRUST

Suppose a smooth piece of lead 50 lb. in weight stands on a smooth horizontal table. Let the area of the surface of the lead in contact with the table be 10 sq. in. Then on 10 sq. in. of the table there is a force of 50 lb. weight spread out equally over that area. On 1 sq. in. of the table there is a force of 5 lb. weight. We say that over the area of 10 sq. in. of the table there is a pressure of 5 lb. weight per sq. in. *Pressure is the force per unit area.*

The total force acting downward on the table due to the weight of the lead is 50 lb. weight. This is called the *total thrust*. It is important to distinguish between the pressure and the total thrust. It is clear from the above example that

Total thrust = Pressure × Area.

LIQUID PRESSURE

We shall first confine our attention to water, the commonest liquid. Carry out the following experiments:

Expt. 248. Suspend a small stone in water by means of a string. Withdraw the stone slowly upwards. There is no hole left. Of course you knew this before you did the experiment, but what does it mean? The water at the side of the stone in its first position flowed inwards to fill the space left by the stone as it was withdrawn to another place.

.Fig. 197. Boyle's experiment on liquid pressure.

The water must have been exerting a pressure inwards on the surface of the stone.

Expt. 249. Punch some holes in the side, and one in the bottom, of a tin can. Close the holes with corks, and fill the can with water. Get a friend to help you and take all the corks out as quickly as possible. Water flows from all the holes no matter where they are situated. Before the corks were removed each must have been exerting a thrust to keep the water from flowing out. Water must exert a pressure outwards on the sides of the vessel containing it.

The following experiment was performed by ROBERT BOYLE. Tubes of the shape shown in Fig. 197 were taken one by one, filled with turpentine, and with the finger closing the top end, lowered into water so that the open ends under water were all at the same depth. The turpentine in the tubes rose above the level of the water, but the turpentine in all three tubes was at the same level. The water at that depth exerted a pressure which forced the less dense turpentine above the level of the water. As the open end under the water in one was facing downwards, in the second upwards, and in the third horizontally, the water at that depth must have exerted an *equal pressure in all directions*, because the level of the turpentine in all the tubes was the same. You feel this pressure when you are in a plunge bath. The water pressure on your chest makes breathing more difficult.

PRESSURE INCREASES WITH DEPTH

Expt. 250. Fill with water a tall cylindrical vessel having outlets at different heights in the sides. The water issues from the bottom outlet with a greater force than from the others (Fig. 198).

The pressure in the water is due to its weight. As water
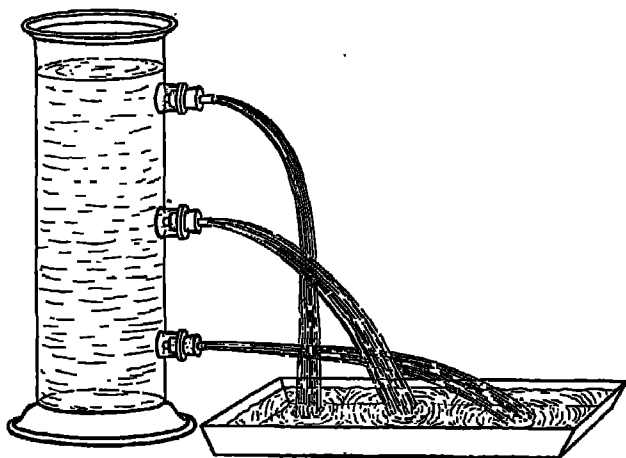
FIG. 198.

weighs 1 gm. per c.c. the pressure on a surface 1 cm. below the surface of the water is 1 gm. per sq. cm. As water is almost incompressible we can assume that the density of the water does not increase with the depth, so that the pressure in the water increases 1 gm. for every increase in depth of 1 cm. In the British System the pressure increases about 1,000 oz. per sq. ft. for every foot increase in depth; or, for every 34 ft. increase in depth there is an increase of pressure of about 14·7 lb. per sq. in.

What we have said about the pressure of water applies generally to liquids. The pressure in water at a depth of h cm. is h gm. per sq. cm. The pressure in a liquid of S.G. d at a depth of h cm. is $h.d$ gm. per sq. cm. These two expressions are really the same because the S.G. of water is 1.

The depth of the liquid (or the difference in depth between two levels) which exerts the pressure is usually called the *head*. In general we can say about any liquid:

$$\text{Pressure} = \text{Head} \times \text{Density}.$$

THE SURFACE OF A LIQUID AT REST IS HORIZONTAL

Expt. 251. Fit up the apparatus as shown in Fig. 199. It consists of a number of glass vessels of various shapes all connected by tubes.
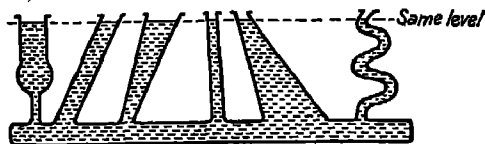
FIG. 199.
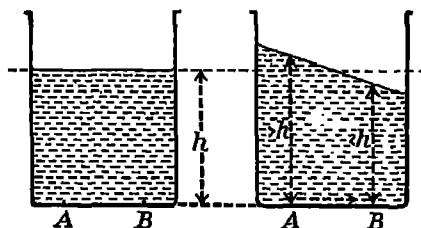
FIG. 200.

Pour water into one of the vessels, and when the water is at rest note the levels of the water in all the vessels. They are all the same height.

Consider water in a beaker (Fig. 200). The pressure at *A* at the bottom is *h* gm. per sq. cm. and the pressure at *B* is the same, as *A* and *B* are at the same depth. Now if the level of the surface of the water is disturbed so that one part of it is higher than another, then, as shown in the Figure, water will flow from *A* to *B* because the pressure at *A* is now greater than that at *B*. The flow will continue until the pressures at *A* and *B* are again the same, and this will be when the surface of the water is horizontal.

This fact has many applications. The level in a vessel of a

liquid not otherwise visible is indicated by an arrangement shown in Fig. 201. It is often used to show the level of oil in large drums, and, in the laboratory, to show how much distilled water there is in the storage jar. When used on a boiler to show the amount of water in it, it is called a *boiler gauge*.

An important application of this fact is related to a town's water supply. We have dealt with this at some length already (p. 13), but we can now understand the principles involved.
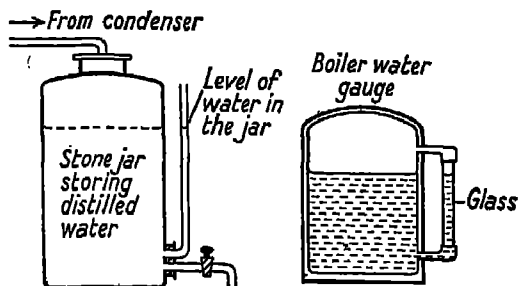
FIG. 201.

The reservoir is situated at a height greater than that of the highest place in the town where water is required. It then flows from the reservoir through pipes to all the places requiring water. It even flows up-hill so long as the hill is lower than the reservoir. It is not sufficient for the reservoir to be just as high as the highest point requiring water, for the water is required to flow out of the tap reasonably quickly, and for this an extra pressure is required. In the lower parts of a town the water comes out of the tap faster (that is at a higher pressure) than in the higher parts. In other words, the taps in the lower parts of the town have a greater head of water than those in the higher parts.

To Find the Densities of Liquids by Means of Fluid Pressure

Expt. 252. To find the density of a liquid which will not mix with water (e.g. an oil). Clamp a U-tube with its long arms, *A* and *B*, vertically and arrange a scale behind it (Fig. 202). Pour some

water in it and then pour the oil down one of the arms. Let Q be a point at the surface of separation, and P a point level with Q in the other arm. Measure on the scale (or otherwise) the height (h cm.) of the free surface of the water above P, and the height (h_1 cm.) of the free surface of the oil above Q.
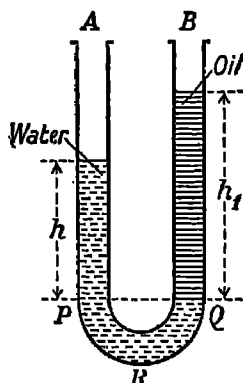
FIG. 202. To find the density of oil.

Pressure at P (due to the weight of water in column A)

$= h \times 1$ gm. per sq. cm.

Pressure at Q (due to the weight of oil in column B)

$= h_1 \times d$ gm. per sq. cm. (where $d =$ density of the oil).

But,

Pressure at $P =$ Pressure at Q (because the lower part of the U-tube PRQ is filled with the same liquid, water).

Hence, $h_1 d = h \times 1$.

Therefore $d = h/h_1$ gm. per c.c.

This method may be used for any two liquids which do not mix. The densities (or specific gravities) are inversely as the heights of the surfaces of the liquids above the surface of separation.

TO FIND THE PRESSURE AT A GAS TAP, USING FLUID PRESSURE

Expt. 253. By means of tubing connect the gas tap with a U-tube containing water (Fig. 203). Turn on the tap. The water rises

FIG. 203.

in arm B and falls in arm A. Measure the height of the surface of the water in B above that in A.
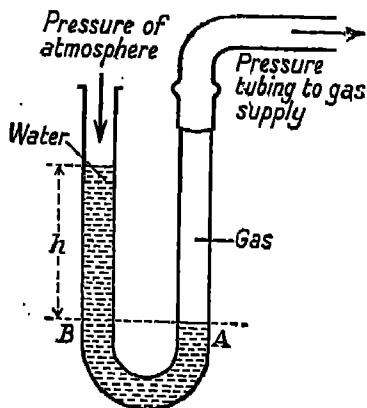
Gas pressure at A = Water pressure at A+Pressure of air,

$\qquad = h \times 1$ gm. per sq. cm.+Pressure of air,

$\qquad = h$ gm. per sq. cm.+Pressure of air.

THE TRANSMISSION OF PRESSURE BY A FLUID

Expt. 254. Cut a small hole in a rubber ball and prick several very much smaller holes in various positions. Fill the ball with water through the larger hole. Covering the latter with the finger, squeeze the ball. Water spurts out from all the small holes. Note the direction of the water spurting out.

The pressure you applied to the water by squeezing the ball was transmitted in all directions at right angles to the surface from all the holes with equal force.

This fact is made use of in the construction of hydraulic machinery. We shall illustrate this by considering the hydraulic press.

The principle of the *hydraulic press* is illustrated in Fig. 204a.

Fig. 204. The hydraulic press.

(It is the same principle as for the Hydraulic Lift.) Consider two cylinders of water A and B, connected by a tube C at the bottom. Let the cross-section of A be 1 sq. in. and that of B 10 sq. in. Apply an extra force of 1 lb. weight vertically downwards on the surface of the water in A. This is transmitted in all directions and there is therefore a vertically upward pressure of 1 lb. weight per sq. in. at the surface of the water in B. The total upward thrust in B is then 10 lb. weight. That is, a force of 1 lb. weight acting downwards at the surface of A is balanced by a force 10 times greater acting upwards at the surface of B.

The details of the hydraulic press are shown in Fig. 204b.

A is the narrow cylinder and B the wide one. A narrow piston travels in *A*, and is operated by a lever, and in this way pressure is communicated to the water in *A*. This is transmitted to the wide piston in *D*, the upward thrust on which is much greater than the thrust communicated to *A*, depending on the ratio of the areas of the cross-sections of the two cylinders. The collar indicated in the figure is to make the joint water-tight.

BOYLE'S LAW

We have dealt with the pressure of a liquid, and we shall now deal with the pressure of a gas. We have already dealt with the pressure of the atmosphere (p. 37) and this should be revised before continuing.
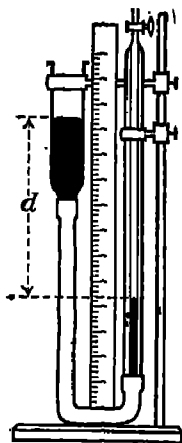
FIG. 205. Boyle's Law.

Expt. 255. To find the relation between the volume and pressure of a gas, the temperature remaining constant. Use the apparatus illustrated in Fig. 205. It consists of a graduated glass tube, the zero mark being at the top. The tube is connected by a strong india-rubber tube to a reservoir, the height of which can be regulated by sliding it on a frame. Pour mercury into the reservoir and close the tap fitted at the top of the graduated tube when the latter is about half full of air. The surfaces of the mercury in both tube and reservoir will now be level because they are both at the pressure of the atmosphere.

Raise the reservoir a little. The mercury in the tube also rises, but not so much as that in the reservoir. The air in the tube is now at a pressure equal to atmospheric pressure plus the pressure due to the difference in levels of the mercury in the tube and the reservoir.

Lower the reservoir until the surface of the mercury is below that in the tube. The air in the tube is now at a pressure less than that of the atmosphere by the pressure due to the difference of the levels of the mercury in the tube and reservoir.

Read the barometer to get the pressure of the atmosphere in millimetres of mercury. Now find the volume of the air in the tube (by reading the graduations) and the corresponding pressure, by measuring on the scale of the frame the difference between the two surfaces

of mercury, and then adding it to (or subtracting it from) the atmospheric pressure. Repeat this for as many positions of the reservoir as possible. Enter your results in a table.

Vol. in c.c. (V)	Height of mercury in		Diff. (d)	Pressure in the tube (P) (Atmos.+d)	P × V
	Tube	Reservoir			

You will notice that the volume multiplied by the pressure is constant. This means that the volume varies inversely as the pressure. That is, if you *double* the pressure you *halve* the volume; or, if you multiply the pressure by four, you divide the volume by four. If a certain amount of air occupies 1 litre at atmospheric pressure, it will occupy 500 c.c. at two atmospheres.

This fact was first discovered by ROBERT BOYLE (1627-91). Boyle used air for his experiments, but other workers have since used many other gases, and in most cases have found that they behave in the same way.

BOYLE'S LAW may be stated as follows: *A volume of a given mass of gas is inversely proportional to the pressure, the temperature remaining constant.*

If some gases, for example carbon dioxide and ammonia, are subjected to much pressure they will liquefy. It has been found that for such gases Boyle's Law does not always apply. All gases can be liquefied by pressure if the temperature is sufficiently low and it has been found generally that if a gas is near the point at which increased pressure will liquefy it, Boyle's Law does not apply.

THE SIPHON

Sometimes it is necessary to transfer a liquid from one vessel to another without pouring. Perhaps the vessel is so large that it cannot be moved. A convenient arrangement for this purpose is the siphon, the working of which depends on the pressure of the air.

Expt. 256. To make a siphon (Fig. 206). *A* and *B* are two vessels. *A* contains water and *B* is empty. *A* stands on a bench higher than *B*. *PQR* is a rubber tube fitted with a clip at the end near *R*. Open the

clip, fill the tube *PQR* with water, and replace the clip. Put the end *P* of the tube in the water in *A*, and the end *R* in the vessel *B*. *Q* is the highest point reached by the tube. Remove the clip. Water flows out of *A* into *B*. Raise the end *R* until it is above the level of the water in *A*. The water no longer flows. Lower the end *R* slowly. As the end falls below the level of the water in *A* the water begins to flow again, and the lower the end *R* the faster is the flow.
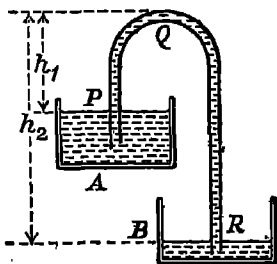
FIG. 206. The principle of the siphon.

Consider what happens. The water flows upwards from *P* to *Q* and then downwards from *Q* to *R*. What causes the water to do this?

Suppose a clip is placed at *Q* while the water is flowing. This stops the flow of the water. At *Q* the head of water due to the arm *PQ* is h_1 (the height of *Q* above the surface of the water in *A*). This tends to pull the water down into *A* but is prevented from doing so by the pressure of the atmosphere on the surface of the water in *A*. The head of water at *Q* due to the arm *QR* is h_2 (the height of *Q* above the surface of the water in *B*). This tends to pull the water down into *B* but is prevented from doing so by the air pressure on the surface of the water in *B*. h_2, however, is greater than h_1, hence there is a greater pull at *Q* towards *R*. When the clip at *Q* is removed the water cannot flow back into both *A* and *B* otherwise there would be a vacuum at *Q*, and this is prevented by the air pressure on the surfaces of the water in *A* and *B*. The water therefore flows towards *R* where there is a greater pull, and continues flowing until the surfaces of the water in *A* and *B* are at the same level.

The siphon principle is sometimes used to transport water over heights. Water in a reservoir situated high in the hills will flow under the action of gravity to its destination so long as the aqueduct does not rise higher than the level of the water in the reservoir. In practice it will not flow quite so high. It is,

however, possible to siphon the water over a slight rise and this is sometimes done. The rise cannot be more than about 30 ft., because the siphon depends on the pressure of the atmosphere. Nevertheless this is quite useful in the transport of water.

The facts about the siphon were known long before they were understood and people were puzzled why water could not be siphoned over high mountains. If it could be siphoned over a slight rise why could it not be siphoned over a greater height? This was not understood until the pressure of the atmosphere was understood.

Normal condition

The siphon principle is often used for emptying household cisterns. It must be remembered that the siphon will not work until the tube is full of water. The pulling of the chain (Fig. 207) fills the tube with water and then the cistern is emptied. The figure gives the detail of the arrangement.

Chain Pulled

PRESSURE GAUGES

In Experiment 253 we found the pressure at a gas tap. The simple arrangement used could be described as a pressure gauge. The same method would not be suitable for finding the pressure at a water tap because this is generally much more than that of the atmosphere, whereas the pressure at a gas tap is generally little more than atmospheric pressure.

Chain released

Fig. 207. The lavatory cistern.

To find the pressure at a water tap a gauge similar to that illustrated in Fig. 208 might be used. It is a closed *manometer*. The end A is fastened very securely to a water tap by means of pressure tubing. When the tap is turned on and the full pressure of the water acts on the surface of the mercury in B, the level of the mercury in C rises, thus increasing the pressure

of the air in C. As this obeys Boyle's Law, the tube C can be graduated to give the pressure for any level of the mercury in C.

For measuring very high pressures the Bourdon Gauge is often used. They may be seen on cylinders in which oxygen and other gases are stored under pressure. They are also commonly used on steam boilers. Fig. 209 illustrates the construction of a Bourdon Gauge. The gas is admitted at A and enters the metal tube B which has an oval cross-section. Under pressure the cross-section tends to become more circular, and this causes the free end of the metal tube to move outwards. The movements are registered on a dial which is graduated to give the pressure of the gas admitted at A. (Re-read Pumps, p. 49.)
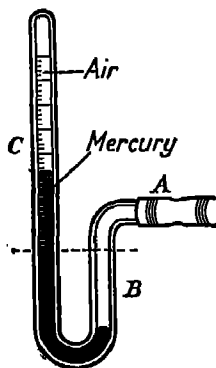
Fig. 208. A closed manometer.

THE EFFECT OF GREAT HEIGHTS AND DEPTHS ON MAN

The pressure of the air is due to the weight of the air. A man on the surface of the earth is at the bottom of a vast and deep ocean of air. The weight of all the air above him causes a pressure of about 15 lb. to the square inch. This acts in all directions, on his body and in his body.

If a man ascends a high mountain he has less air above him, and therefore the pressure of the air on him is less. As the height increases the density of the air decreases until at a very great height there is no pressure due to the air because there is no air. This happens at a much greater height than that of the highest mountain.

Consider a man climbing a mountain. For the first few thousand feet he will probably not notice any change in the air around him, but as he continues to ascend he finds it increasingly difficult to breathe, and he is much more easily fatigued. When he fills his lungs with air, the volume taken in is roughly that taken in at sea-level; but on the mountain the air is at a reduced pressure. Consequently the mass of air taken in on the

mountain is less than that taken in at sea-level. This means that the mass of oxygen taken in is less. For this reason the man finds himself short of oxygen in spite of the fact that there is an enormous supply around him. Respiration becomes
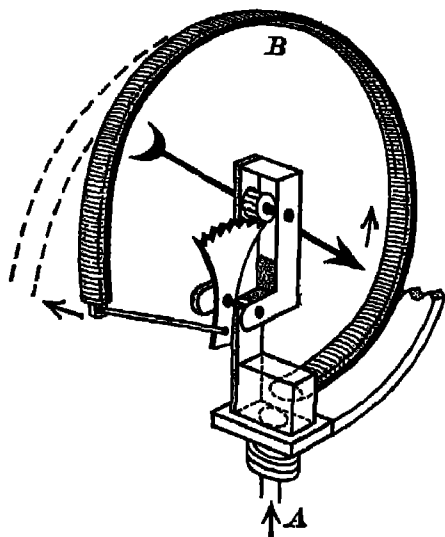
FIG. 209. The Bourdon Gauge.

difficult, and his energy, which depends on his respiration, decreases.

Climbers attempting the ascent of very high mountains often wait for days at various heights to get used to the decreased pressure before proceeding to heights where the pressure is still less. Expert climbers seem to adapt themselves to great heights but even their power of accommodation is limited. Some carry oxygen with them, but this is not all gain as the extra weight of the oxygen apparatus causes more fatigue. A man ascending in an aeroplane is not in the same difficulty because his plane carries his oxygen, and in any case there is not the same demand on his energy.

The pressure of the air affects divers. Consider a diver 100 ft.

below the surface of the water. The pressure at that depth is about 60 lb. per sq. in. Air must be pumped to the diver at this pressure. This is about four times the pressure that man normally experiences yet the diver experiences little discomfort. He is able to work as well as the resistance of the water will allow.

It is found, however, that if a diver, after being at that depth, returns to the surface quickly, he suffers pain in the limbs, may become unconscious, and may even die. If the return to the surface is gradual no such discomfort is experienced. Let us see the reason for this.

If a certain amount of gas dissolves in water at atmospheric pressure, then more gas dissolves in the same amount of water if the pressure is increased. This may be seen on opening a bottle of mineral water. When the stopper is removed bubbles of gas (carbon dioxide) leave the liquid in which they have been dissolved under pressure. The opening of the stopper reduces the pressure. When a diver is at a great depth under water he is breathing air at a great pressure and more than a normal quantity of the gases of the air dissolve in the blood and in the other liquids of the body. This seems to cause no discomfort. When the diver comes rapidly to the surface he is in air at a greatly reduced pressure, and the gases dissolved in the liquids come out of solution so quickly that bubbles of gas are formed in the tissues. This causes the pain in the limbs. If the bubbles form in places like the brain or blood-vessels they may cause paralysis and death. When a diver is accidentally brought up too quickly and shows symptoms of this kind he is enclosed in a pressure chamber where the pressure of the air is gradually reduced. In the construction of tubes and tunnels the men work in air at a very high pressure. An air-locked chamber is provided for them where they sit while the pressure is slowly reduced to that of the normal atmosphere. They can then leave the chamber.

Much of a diver's time is spent, not in working, but in coming to the surface. For this reason diving at great depths in an ordinary diving suit is impracticable. Another method is to send the diver down in a steel shell supplied with air from

the surface and connected by telephone to the ship. From this shell the diver can see what needs doing, but he can do nothing himself except direct operations. The men on the ship manipulate the cranes and grappling hooks directed by the diver in the shell.

QUESTIONS

1. A cylindrical tank of diameter 14 in. contains water to a depth of 3 ft. What is (a) the total thrust on the bottom of the tank, (b) the pressure at the bottom, and (c) the pressure on the side of the tank 1 ft. from the bottom? What would be the answers to (a), (b), and (c) if the water were replaced by oil of S.G. 0·85? (1 cu. ft. of water weighs 62·3 lb.)

2. An inverted U-tube with one limb, A, in water and the other limb, B, in oil has air drawn from it so that water is drawn up A and oil up B. The level of the water in limb A is 25 cm. and of the oil in limb B is 30 cm. What is the S.G. of the oil? Explain the principle on which you work.

3. Explain the principle and uses of the siphon.

4. Describe, with the aid of a sketch, a common type of pressure gauge. What is its use and how does it perform its function?

5. Write an essay on 'The effects of great heights and depths on Man'.

6. If a small stone is thrown overboard into the deepest part of the ocean will it ever reach the bottom of the ocean? Give reasons for your answer.

XXIV

HEAT. EXPANSION

HEAT

WE are all familiar with the idea of heat. We feel the warmth of the sun or a fire. We know that we are warmer than our surroundings except perhaps on a very hot day. We notice heat produced in many different ways. The importance of heat to us and other animals is obvious, and we know that even plants stop growing, or grow very slowly, in cold weather. Heat is closely connected with our everyday life, and all matter, living and non-living, is affected by it. It is natural then for us to ask what is heat, and how does it affect matter?

THE CONDITIONS UNDER WHICH HEAT IS PRODUCED

We know from our ordinary experience that an object becomes warmer when the sun shines on it. This fact may be illustrated by allowing the rays of the sun to pass through a 'burning-glass' (a convex lens) which brings many rays to one point (Ch. XXXIII). If the rays are concentrated on one spot on a piece of paper, the latter will soon begin to char and will perhaps take fire.

When a substance burns, heat is produced. If the combustion is rapid the heat is obvious, but even in slow combustion heat is produced though it is often difficult to detect without special precautions. In all oxidation, vigorous or slow, heat is produced. In many other chemical actions heat is produced, sometimes when substances combine, and sometimes when they decompose. Every chemical action is accompanied by a heat change; heat is either given out or taken in. This heat change is sometimes the only outward sign that a chemical action has taken place. Heat is a form of energy and may be transformed into chemical energy. On the other hand, chemical energy may be transformed into heat energy.

When two pieces of wood are rubbed together vigorously

they become warmer. If suitable wood is chosen and the friction is very great the wood may take fire. Similarly if a piece of iron is well hammered both the metal and the hammer become warmer. In both cases, the rubbing and the hammering, the mechanical energy is transformed into heat energy. In a steam engine the reverse process takes place, heat energy being transformed into mechanical energy.

When an electric current is passed through a conductor heat is produced. The electrical energy is transformed into heat energy. In the case of a steam engine driving a dynamo, heat energy is transformed into electrical energy.

The interior of the earth is very hot. We know this because the farther we descend into the earth the hotter it becomes. Hot springs and volcanoes also suggest that the interior of the earth is very hot. This internal heat is not a convenient practical source of heat and its existence is often forgotten.

The chief source of heat energy on the earth is the sun (Ch. XLIII).

HEAT IS A FORM OF ENERGY

In our study of the composition of matter (p. 158) we saw that matter is composed of molecules. These molecules are not at rest but are continually vibrating. The molecules are prevented from moving away by a force called *cohesion*. In a solid this force is strong; in a liquid it is not so strong as in a solid; in a gas it is so small as to be negligible.

When a body is heated the molecules move more quickly, and the hotter a body becomes the faster the molecules move. As the body becomes colder the speed of the molecules decreases until a point is reached when the molecules are at rest. The body is then as cold as it can be. This condition has never yet been reached in practice, but it has been very nearly attained.

When one form of energy is transformed into heat energy it really causes an increased kinetic energy of the molecules of the body. In the case of a piece of metal being hammered, some of the kinetic energy of the moving head is transformed into the kinetic energy of the molecules of the metal and the hammer

head, and this appears to our senses as heat. When the sun warms a body, some of the heat of the sun is transformed into kinetic energy in the molecules of the body.

HEAT AND HOTNESS

It is necessary for a clear understanding of the subject to distinguish carefully between the amount of heat a body has, and the hotness of the body.

When solids or liquids receive heat (without changing their state) they become hotter, and when they lose heat they become colder. If there are two lumps of iron, both of the same hotness but one as big as a cricket ball and the other as big as a football, then the larger one has more heat than the smaller, although it is no hotter than the smaller one. When we say that a body is hot, we really mean that it is hot compared with the ordinary things around us which we handle. We say that ice is cold, but compared with liquid air it is very hot. A kettle of liquid air put on ice boils vigorously. This shows that even cold things like ice contain heat. We must return to this subject later when we have learned something about heat, but it is necessary even at this stage to distinguish between heat and hotness (p. 423).

The hotness of a body when compared with a standard hotness is called *temperature* and we shall use this term in future although we have said nothing yet about the standard of hotness.

We have said that when heat is given to a body it may become hotter, but this is only one of the effects of heat on a body. Bodies expand on heating, some change their state, others decompose, and there are sometimes electrical effects.

SOLIDS EXPAND ON HEATING

Expt. 257. Fasten a long steel knitting needle to a clamp, as shown in Fig. 210, so that it cannot increase in length without bending. Now heat the needle with a bunsen burner. The needle bends. Allow the needle to cool. It becomes straight again.

Expt. 258. Obtain a brass ball which will just pass through a metal ring. Heat the ball and then place it on the ring. It will not pass through at first but later it does so. On heating, the brass ball expands,

on cooling it contracts. While cooling on the ring, some of the heat passes to the ring which becomes hotter and expands (Fig. 210).

Solids expand on heating and contract on cooling. A few, like india-rubber, do not. Some alloys (mixtures, or perhaps peculiar compounds of different metals) have been made which expand only to a very small extent (p. 265).
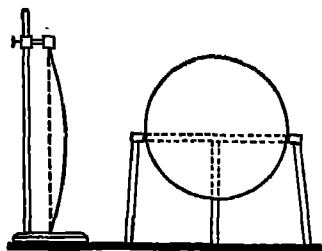
Fig. 210. The expansion of solids on heating,
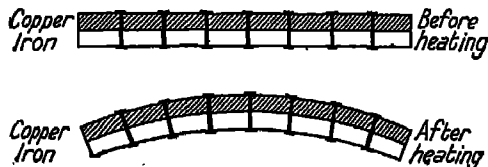
Fig. 211.

Expt. 259. Take a bar which has been made by riveting together two flat strips of different metals, e.g. copper and iron. Note that the bar is straight. Now heat the bar equally on both sides with the bunsen flame. The bar bends. As the bar cools it becomes straight again (Fig. 211). The metals expand unequally. As the metals were riveted together, the only way in which one metal could expand more than the other was for the compound bar to bend. The outside curve is longer than the inside one. The figure makes this clear.

Different solids expand different amounts when heated to the same temperature.

The force exerted by solids on cooling can be shown by the following experiment.

Expt. 260. Use the apparatus shown in Fig. 212. Heat the iron bar *A* very strongly and then fix the cast iron bolt *B* in position, afterwards tightening the screw *C* as much as possible. Allow the bar to cool. As it does so it contracts and the bolt is broken.

This great force exerted by contracting bodies is often put to practical use.

SOME EFFECTS OF THE EXPANSION AND CONTRACTION OF SOLIDS

Railway lines expand in the summer and contract in the winter. Space is left between each length of line to allow for this. At the same time it is necessary for the lines to be stable. The next time you get an opportunity of *safely* examining the joinings of two lengths of line, do so. Fig. 213 will help you to understand the method. No space is allowed between the lengths of tram-lines. These, unlike railway lines, are sunk in the road and are not subject to the same extremes of temperature. In addition to this, any strain is taken by the material in which they are sunk, usually concrete.

Metal bridges are longer in summer than in winter. Allowance is made for this, otherwise strains are set up in the bridge structure which might in time make the bridge unsafe.

The moving parts of machines become hot after a time owing to friction, even if there is no other cause. They expand, and allowance has to be made for this. The piston of the internal-combustion engine is made to fit in the cylinder easily. When it gets hot owing to the heat of the explosions and the friction, it expands and fits more tightly in the cylinder. Until an internal-combustion engine is warm it is not running at its best because the piston is too small. If the oil supply fails, the engine becomes too hot and the piston expands more than it should. This makes matters worse, for it increases the friction. If the engine continues running under these conditions it will 'seize', that is, the piston will get stuck in the cylinder (p. 484).

Pipes which carry hot water expand. The different lengths of pipe fit into each other and the joint is packed with asbestos. This arrangement allows for expansion (Fig. 215).

A clock keeps good time only if the length of the pendulum is constant. Unless some compensation is made a clock loses in
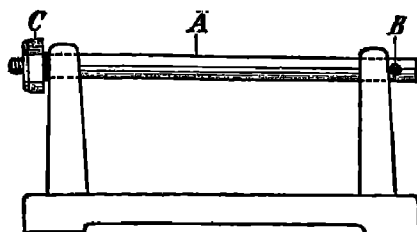
FIG. 212. The great force exerted by a contracting bar of iron.

Fish-plate   *Bolt head*

Fish-plate with
oval bolt holes
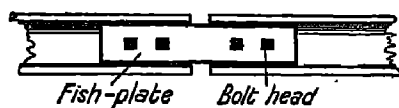
FIG. 213. How railway lines are joined.

Roller bearing   *Fixed*
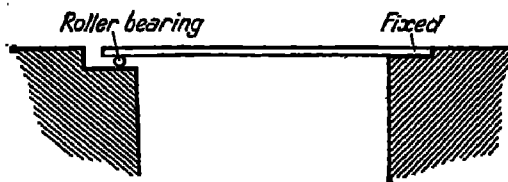
FIG. 214. The construction of a bridge.

Asbestos
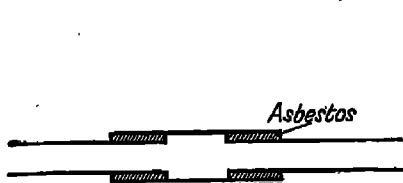
FIG. 215. The joining of hot water pipes.

Wood (expands
downwards)

Brass (expands
upwards.)
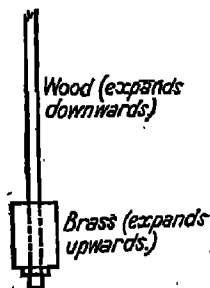
FIG. 216. A pendulum.

summer and gains in winter. To prevent this there is an arrangement by which the bob of the pendulum expands upwards while the pendulum itself expands downwards (p. 324).

If hot water is poured into a thick tumbler there is a possibility of the tumbler cracking. This is due to the unequal
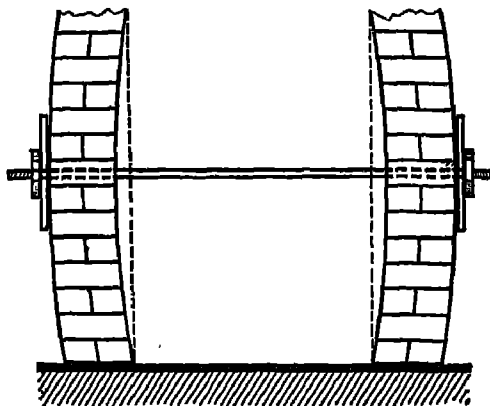
FIG. 217. Old walls brought back to the vertical.

expansion of the glass, the inside expanding more than the outside.

In those parts of the world which suffer extreme temperatures (deserts), the rocks are subjected to alternate expansion and contraction. This causes them to disintegrate into sand, but this is not the only cause of disintegration (Ch. XLVII).

APPLICATIONS OF EXPANSION AND CONTRACTION

Sometimes the walls of a building have a tendency to fall outwards. To prevent this an iron rod is passed between the two walls (Fig. 217). The rod is strongly heated in the middle and while it is still hot the wall plates are screwed up tightly. On cooling the rod contracts, drawing the walls together. This is an application of Experiment 260, but in this case the bolt (wall plate) does not break.

A wheelwright fits an iron tyre on to a wooden wheel in the following way: The tyre is made so that when cold it is smaller in circumference than the wooden wheel. The tyre is then heated all round in a large fire. While hot it is fitted to the wheel and cooled quickly by revolving the wheel in water. As the tyre contracts it fits tightly to the wheel, binding it securely.

The metal platinum and glass expand equal amounts on being raised to the same temperature. When a hot platinum wire is fitted into glass made soft by heating, the wire remains firmly fixed in the glass on cooling. This would not be so if the wire was of copper, for copper and glass do not expand equally. Porcelain also expands to the same extent as platinum; for this reason platinum pins are often used by dentists in making artificial teeth.

FIG. 218. Water expands on heating.

Other applications of expansion and contraction are dealt with later, e.g. the metallic thermometer (p. 431) and the incubator (p. 431).

LIQUIDS, AS A RULE, EXPAND ON HEATING

In this respect they are like solids. The fact may be illustrated by taking water as the liquid.

Expt. 261. Fit a small flask with a one-holed cork through which passes a glass tube with a small bore (Fig. 218). Fill the flask with water coloured with red ink. When the cork is replaced some of the coloured water should rise in the tube. Heat the flask gently. Note the rise of the water in the tube as the heating proceeds. This means that the volume of the water in the flask increases on heating. Actually the volume of the flask also increases as the glass expands on heating. The increase in volume of the flask, however, is very much smaller than that of the water. The expansion of the water noted is called the *apparent expansion*. The *real* expansion is greater.

Different Liquids Expand Different Amounts when Raised to the Same Temperature

Expt. 262. Arrange three specimen tubes of equal volume and fitted with the same kind of tube as shown in Fig. 219. Fill A with water, B with mercury, and C with alcohol. Arrange that the heights of all three are the same at the temperature of the room. Put the three together in the same bath of warm water, and leave until
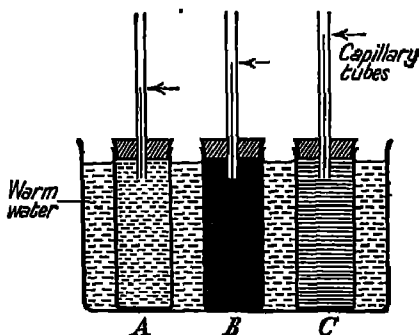
Fig. 219. The expansion of liquids on heating.

the heights of the columns of liquids are steady. Note that all three reach different heights. Which rises the highest and which the least?

Each was raised to the same temperature, as they were all put into the same bath of water. The expansions were different.

The expansions of the three liquids in this experiment were visible to the naked eye. The expansions of the solids in Experiment 259 were not visible without some device to make it obvious. Liquids, as a rule, expand much more than solids for the same rise in temperature.

GASES EXPAND ON HEATING

We shall use air to illustrate this.

Expt. 263. Fit a one-holed cork to an 'empty' flask, that is, a flask full of air. Through the hole fit a long glass tube of small bore con-

taining coloured water in about 1 in. of its length (Fig. 220). Run cold water on the outside of the flask. What happens to the coloured water index? Now hold the flask in your warm hands. Note what happens to the index. The air in the flask expands very much even with such slight warming as the hand gives. For comparison use a flask of the same size full of water and repeat the experiment. The difference in temperature between the cold water and the hand is not sufficient to show any marked movement of the index.

Air expands very much more than water, and generally it can be said that gases expand much more than liquids for the same rise in temperature.

The expansion of gases is a very important subject and we must return to it for a more detailed study when we have learned more about temperature (p. 437).

TEMPERATURE AND HEAT

We have already said something about heat and hotness (p. 416). It is now necessary to say more about this. We can tell to some extent by feeling a body, or by going near to it, whether it is hot or cold. Some parts of our body are more sensitive than others in this respect. When we wish to 'feel' if a body is hot, we often place it near our cheek. We place our hands in water to see if it is hot enough for our purpose, and for most ordinary purposes this method of testing the hotness of a body is good enough. For scientific and for many practical purposes this is not good enough. One reason for this is that our skin is not sensitive enough to appreciate small differences in hotness. It may be all right for telling us whether a body is hot or very hot, but often we wish to know which of two bodies is the hotter when there is very little difference in their degree of hotness. Another reason why we cannot depend on our sensations of hotness is that the sensation depends on the preceding sensation. An experiment will make this clear.

Coloured water index

Air
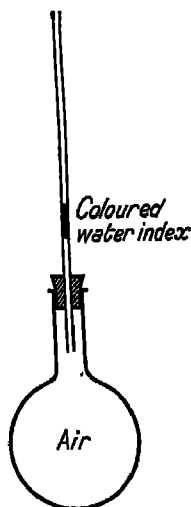
FIG. 220.
The expansion of air.

Expt. 264. To show that we cannot rely on our sensations of hotness. Place three dishes of water side by side. Let *A* contain hot water, *B* cooler water, and *C* cold water. Put the left hand in *A* and the right hand in *C*. The left hand will, of course, feel hot and the right hand cold. Remove the hands and place both of them in *B*. Both hands are now in water of the same hotness but they do not feel the same. The right hand feels hot and the left cold.

In this experiment our left hand tells us that whereas it was hot it is now cooler; the right hand indicates that whereas it was cold it is now warmer. This, of course, is valuable information, but it indicates that we cannot rely too much on the hand as a test of the hotness of a body.

For this reason instruments have been invented to measure hotness. They are called *thermometers*. The hotness of a body when measured by a thermometer is called its *temperature*. It must be clearly understood that thermometers measure hotness and not the amount of heat.

THERMOMETERS

These are now common objects of ordinary life. We keep one to tell if the baby's bath is just right, another to test the baby's milk, a third for the greenhouse. One is kept in the schoolroom as a check on over- or under-heating, and the doctor sometimes puts one under our tongue when we are ill. These thermometers are of many shapes and sizes, but if you examine them you will find that they nearly all have one thing in common. They register temperature by the expansion of a liquid as it gets hotter and by the contraction of the liquid as it gets colder. You should examine as many different thermometers as possible.

In England the thermometers in general use are called *Fahrenheit* thermometers and indicate the temperature at which water boils as 212°. This is called the *boiling-point* of water. The *freezing-point* of water is marked at 32°.

FIG. 221. Centigrade thermometer.

Sometimes there is another mark at the temperature of the human body. This is called Blood Heat, a very bad name for it, because a thermometer does not measure heat but hotness. We shall consider first another kind of thermometer which is in general use in scientific work all over the world. It is called the *Centigrade* thermometer.

THE CENTIGRADE THERMOMETER

Examine a Centigrade thermometer (Fig. 221). It consists of a very fine tube connected with a comparatively large bulb. The liquid is (usually) mercury. Hold the bulb in your fingers and watch the rise of the mercury in the tube. Now remove the fingers. The mercury falls.

In making a thermometer it is necessary to have some standard. It is found that the temperature of a mixture of pure ice and pure water is always the same. The temperature of pure boiling water (at the same pressure) is always the same. These two temperatures are taken as fixed points. The *freezing-point of water* (or the melting-point of ice) is taken as $0°$ in the Centigrade scale. The *boiling-point of water* (at 760 mm. pressure) is called $100°$ on the same scale.

Expt. 265. To test the lower fixed point of a thermometer. Break up some pure ice into small pieces and place them in a filter funnel (Fig. 222). Embed the bulb of a thermometer in the ice and pour some distilled water over the ice. Watch the fall of the mercury in the thermometer. If the thermometer is a good one the mercury will fall to $0°$ and then stay there.

FIG. 222. To test the lower fixed point of a thermometer.

Expt. 266. To test the higher fixed point of a thermometer. Boil some distilled water in a flask in the neck of which is suspended a thermometer. Let the water boil vigorously so that the steam is surrounding the bulb. Under such conditions the water and the steam are at the same temperature if the water is pure. The thermometer will register $100°$. (It may not do so, because with this simple apparatus it is difficult to keep the steam from condensing.)

The experiment may be more accurately performed by using a hypso-meter (Fig. 223). Most of the stem and the bulb are immersed in the steam which then forms a steam jacket round the apparatus. The bulb must be out of reach of splashes and a water manometer is fixed at the side to ensure that the pressure inside is the same as that outside. This is obtained by reading the barometer. The vessel is
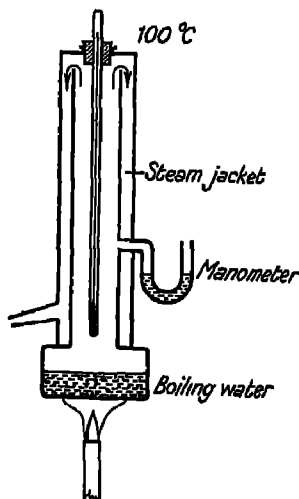
FIG. 223. To test the upper
fixed point of a thermometer.

made of copper. In the standardization of good thermometers the pressure of the steam must be 760 mm.

THE EFFECT OF A DISSOLVED SUBSTANCE ON THE
BOILING-POINT AND FREEZING-POINT OF WATER

Expt. 267. To show the effect of dissolved salt on the boiling-point of water. Place the bulb of a thermometer in some boiling distilled water in a beaker. Note the temperature registered. (It may not be exactly 100° C.) Now add some salt to the water in the beaker. Note the temperature registered when the solution is boiling. The boiling-point of the solution is higher than that of the water. A dissolved substance always raises the boiling-point.

Expt. 268. What is the temperature of the steam which comes from a solution of salt in water? Repeat Experiment 266 and note the temperature registered. Now add some salt to the water and note the temperature of the steam. Be sure that the bulb of the thermometer is not splashed by the solution. (Why?) The temperature of the steam is not altered.

Expt. 269. To show the effect of salt added to water on the freezing-point. Repeat Experiment 265 but this time mix salt with the ice. What temperature is registered? Dissolved substances lower the freezing-point.

There is a common application of this fact. Salt is often thrown on ice-covered roads to thaw the ice. A solution of salt in water will not freeze at 0° C. Of course, if the temperature falls low enough, the solution will freeze. In England the temperature seldom falls so low that salt-water freezes.

COMPARISON OF THE CENTIGRADE AND
FAHRENHEIT SCALES

Fig. 224 shows the relation between the two scales. It is often necessary to change from one scale to the other. This is a matter of simple arithmetic. Between the freezing-point and the boiling-point on the Centigrade scale there are 100 degrees. Between the same fixed points on the Fahrenheit scale there are (212 — 32) that is 180 degrees.

FIG. 224.
Comparison of Fahrenheit and Centigrade thermometers.

Hence   1 C. degree = $1\frac{4}{5}$ F. degrees;

or,   1 F. ,, = $\frac{5}{9}$ C. ,,

Also   32° F. is the same temperature as 0° C.

If C.° on the Centigrade scale is to represent the same temperature as F.° on the Fahrenheit scale, then,

$$\frac{C.}{5} = \frac{F. - 32}{9}.$$

When a number of readings have to be changed from one scale to another it is convenient to make a graph. If this is done to a suitable scale the result can be made as accurate as we please. Fig. 225 illustrates such a graph.
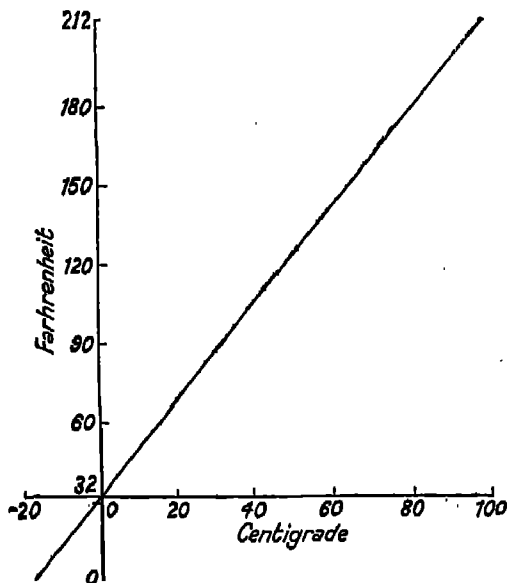
Fig. 225. Graph for conversion of temperature from one scale to another.

·THERMOMETERS FOR SPECIAL PURPOSES

It is convenient to have thermometers specially constructed to serve particular purposes. These may be graduated with the Centigrade or the Fahrenheit scale. The following examples are given but there are many others:

1. The clinical thermometer

This is used by doctors and nurses to take the temperature of a patient's body. In normal conditions of health, the temperature of the human body is about 98·4° F. If it varies much

from this temperature the patient's life may be in danger. There is, therefore, no need for the clinical thermometer to be graduated from 32° F. to 212° F. It is usually graduated from 95° F. to 110° F. Each degree is also subdivided so that the temperature of the body can be read accurately to one-fifth of a degree. It is a mercury thermometer constructed as shown in Fig. 226. The bulb is comparatively large, and the bore very thin. There is a constriction in the bore just above the bulb. The bulb is put in the patient's mouth or in his arm-pit and left there for a short time. It is then removed, but the mercury does not run back into the bulb, because it is prevented from doing so by the constriction. The doctor is thus able to read the temperature accurately although the thermometer is not now in contact with the patient. Before taking the tempera-ture of another patient, the thermometer must be shaken to force the mercury back past the constriction to the bulb. Clinical thermometers must not be washed in hot water because the high temperature would cause such an expansion of the mercury that the instrument would be damaged.
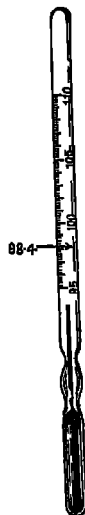
Fig. 226. Clinical thermo-meter.

2. The maximum and minimum thermometer

This is used chiefly for taking temperature observa-tions out of doors. Usually the highest and lowest temperatures during the twenty-four hours preceding the observation are required. There are many different forms of this instrument, but all depend on one principle: something must remain fixed at the highest and lowest points reached during the period, no matter what subsequent changes of temperature take place.

One form of maximum and minimum thermometer is shown in Fig. 227. It consists of two thermometers, the bores of which remain horizontal. The maximum thermometer contains mercury and, like the clinical thermometer, has a constriction in the bore just above the bulb. When the temperature rises, the mercury is forced from the bulb past the constriction, but when the temperature falls, the mercury is prevented from

returning to the bulb by the constriction. Thus the highest temperature reached is indicated by the end of the thread of mercury.

The minimum thermometer contains alcohol instead of mercury. In the bore there is a dumb-bell-shaped index of glass or steel. As the temperature rises, the alcohol flows past the indicator without altering its position, but when the tem-
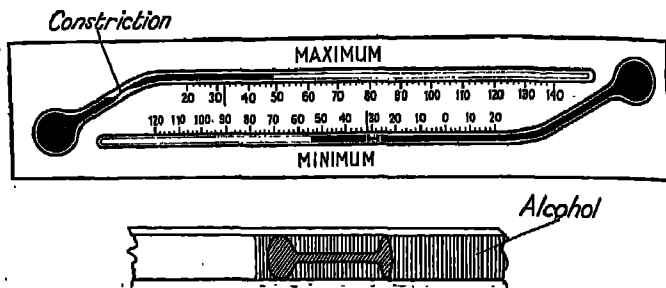
FIG. 227. One form of maximum and minimum thermometer.

perature falls, the index is drawn back by the meniscus in the bore.

Alcohol is used instead of mercury because the former wets steel and glass and can thus flow past it. Mercury does not wet these substances and it would push the indicator along with it.

3. The metallic thermometer

Some motor-car engines are fitted with a device for indicating roughly the temperature of the engine. It shows whether the engine is cool, hot, or normal. The device is a rough metallic thermometer and consists of two strips of different metals, e.g. iron and brass, bound together. As the temperature increases the metals expand unequally and bend in one direction. This causes a ratchet to turn a small cogwheel which is attached to a pointer (Fig. 228).

THE THERMOSTAT

Sometimes it is necessary to keep a liquid or a volume of gas at one temperature. There are many ways of doing this, some of them very elaborate. Any apparatus which does this is usually called a *thermostat*. The principle of the thermostat is
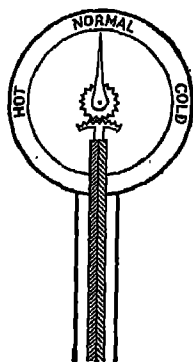
FIG. 228. A metallic
thermometer.
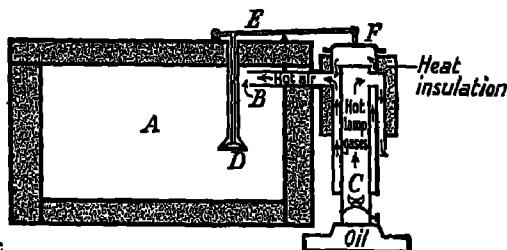
FIG. 229. An incubator.

illustrated by the way the air in an incubator is kept at a fairly constant temperature (Fig. 229).

An incubator consists essentially of a box, *A*, where the eggs to be hatched are placed. Hot air enters the box through the hole *B* from around the lamp *C*. In the box is a strip of iron bound at both ends to a strip of aluminium, *D*. As the temperature of the air in the box increases both the iron and the aluminium expand, but the latter expands more than the iron. The aluminium therefore bends upwards like a bow and operates a lever *E* which opens the top of the damper *F*. This allows the hot air to escape to the outside air. The temperature of the air in the box falls, and as it does so the aluminium contracts and the lever operates in the reverse direction, thus closing the top of the lamp. The hot air now enters the box again and the cycle of operations continues. The result is that the incubator and the eggs in it are kept within a very small range of temperature.

This arrangement is sometimes modified to employ an electric heater instead of an oil lamp. The temperature range in an incubator need not be very small and this simple arrangement is good enough for the purpose. Some thermostats used in the laboratory can keep a volume of air or water at a temperature which is, for all practical purposes, constant. Examine one if you get the opportunity and try to discover how it works and what range of temperature is kept.

A similar device involving the unequal expansion of different metals is used in the construction of some stoves and ovens, both electric and gas. The cook can arrange for the oven to remain at a particular temperature just by operating a small lever. Some fire and frost alarms are also constructed on the same principle.

WHAT IS THE BEST LIQUID TO USE IN A THERMOMETER?

In all the thermometers (except metallic thermometers) we have mentioned so far, either mercury or alcohol was used. Why is mercury so often used? Water, above 4° C., expands on heating and yet this is seldom used as a liquid in thermometers. What are the reasons for choosing one liquid rather than another?

A mercury thermometer can be used over a wide range of temperature, as the boiling-point of mercury is 357° C. and its freezing-point is —39° C. An alcohol thermometer cannot be used for high temperatures as the boiling-point of alcohol is 78° C., but it is very useful for measuring low temperatures as its freezing-point is —114° C., very much below that of mercury. Mercury does not wet the glass tube and alcohol does. Hence after a quick fall of temperature the alcohol thermometer gives too low a reading. Mercury is a good conductor of heat while alcohol is not so good. On the other hand, alcohol expands much more than mercury for the same rise in temperature. Mercury is easily seen in the tube and a much thinner thread suffices, but if alcohol is coloured it can be seen as easily as mercury. Mercury also expands regularly but alcohol does not.

It is clear now why water is not commonly used in thermometers. Its range is small. (F.P. 0° C., B.P. 100° C.) It wets

the tube. It is a bad conductor of heat, and does not expand so much as mercury for the same rise in temperature. Below 4° C. it is unreliable as we shall see later. Also 1 c.c. of water requires more than twice as much heat as 1 c.c. of mercury to raise its temperature 1° C.

THE FREEZING OF WATER

Expt. 270. To show that water expands on freezing. Fill a small bottle with water and close it with a screw cap. Immerse it in a mixture of salt and ice. Leave it for about a quarter of an hour. A longer time may be necessary if the bottle is large. Examine the bottle. It will be found broken.

The water expanded as it froze. As the bottle was already full of water, there was no room for all the ice formed and so the pressure broke the bottle.

Lead water-pipes sometimes burst during frosty weather because the water in them freezes. The burst is often not noticed until the thaw sets in and the water begins to run out. One way of preventing the water-pipes from bursting is to allow the water to flow slightly all the time. Running water does not freeze so easily as still water. This method, however, is very wasteful and is not to be recommended. A better method is to cover the part of the pipe exposed to the weather with some non-conducting material such as straw, hay, or layers of paper (p. 460). In modern buildings it is usual to arrange that all water-pipes are protected from the weather as far as possible.

Another consequence of the expansion of water on freezing is that ice floats. The density of ice is less than that of water, therefore, the ice must float. The density of ice, however, is not very much less than that of water, and so a large part of the ice is submerged. A floating iceberg may appear enormous but there is nine times as much under water as above it.

All liquids do not expand on freezing. Paraffin wax does not. For this reason solid wax does not float on liquid wax but sinks to the bottom.

THE BEHAVIOUR OF WATER AS IT COOLS FROM 10° C. TO 0° C.

Expt. 271. To investigate the behaviour of water as it cools.
Fit up the apparatus shown in Fig. 230. The flask is full of coloured water. Note the level of the column of water in the thin tube, and the temperature. Immerse the flask in a mixture of crushed ice and salt. As the temperature falls the level of the water in the tube falls. Note that somewhere about 4° C. the level of the water in the tube stops falling and after that, as the temperature falls, the level rises.

Level of coloured water in the tube

Thermometer

Coloured water

A mixture of ice and common salt

FIG. 230.

As water is cooled down to 4° C. it contracts and becomes denser. Below that temperature it expands and thus becomes less dense. The *maximum density of water* is at 4° C. We have seen that at 0° C., when water freezes, the expansion is still greater.

In the above experiment, owing to the contraction of the glass the apparent minimum will probably not be at 4° C. but more likely at 6° C.; using a silica flask would give a result nearer 4° C. Hope's experiment is usually used to illustrate the fact that the maximum density of water is at 4° C.

Water

Ice and salt

FIG. 231. Hope's experiment.

Expt. 272. Hope's Experiment. The apparatus is as shown in Fig. 231. Distilled water is placed in the jar and the middle surrounded with a mixture of ice and salt. Readings of both thermometers are taken at frequent intervals. The lower thermometer falls to 4° C. and then remains stationary. The upper thermometer falls to 0° C. and the surface of the water freezes.

The explanation is as follows: The water in the middle becomes colder, contracts, and therefore becomes denser. It sinks to the bottom, the less dense water rising to take its place. At 4° C. any cooling causes an expansion, with a consequent decrease in density. The water then rises from the middle to the top, warmer but denser water sinking to take its place. When this water reaches the middle it is cooled to 4° C. (or below that) and in turn it rises to the top. This continues until all the water at the top is at 0° C. when it freezes on losing more heat.

It is a consequence of the fact that the maximum density of water is at 4° C. that ponds freeze at the top and not from the bottom upwards. Consider a pond being cooled by the air above it. The top layer becomes colder, contracts, and being denser than the water below, sinks to the bottom, warmer water from below taking its place. This continues until all the water in the pond is at 4° C. The top layer now becomes cooler, but in doing so expands and becomes less dense. It cannot therefore sink but remains at the top until the temperature reaches 0° C., when the top layer freezes. If the maximum density of water was not at a temperature above its freezing-point ponds would freeze from the bottom upwards, and if during a hard winter the whole pond were thus frozen it might not thaw before the summer was over. When we have studied the transference of heat we shall return to this subject. It is almost certain that life, as we know it, could not exist on this planet unless the maximum density of water was at a temperature above its freezing-point. It is strange that water is exceptional among liquids in having this property.
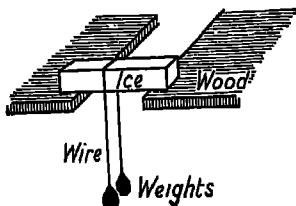
Fig. 232. Regelation.

THE EFFECT OF PRESSURE ON THE MELTING-POINT OF ICE

Expt. 273. Support a block of ice as shown in Fig. 232. Pass over it a thin copper wire with a weight attached to each end. Note that the wire gradually passes through the block of ice but the latter is not cut in two.

The wire exerts a pressure on the ice and this lowers the melting-point. The ice melts under the wire and then freezes above it. This continues until the wire has passed right through the ice. The process of melting and then freezing again is called *regelation*.

When snow is pressed together to make a snowball the snow melts and then freezes again, thus making a firm snowball. In very cold weather it is impossible to make snowballs. This is because the effect of the pressure is small and a very great pressure is necessary to cause melting at low temperatures.

THE BOILING OF WATER UNDER REDUCED PRESSURE

Expt. 274. To investigate the boiling of water under reduced pressure. Boil some water vigorously in a strong round-bottomed flask. Remove it from the flame and immediately, while the steam is still issuing from the mouth of the flask, insert a well-fitting rubber stopper through which passes a thermometer. Pour cold water on the outside of the flask to cool it. Note the vigorous boiling which takes place when the cold water touches the flask. At the same time read the temperature from the thermometer. (The rubber stopper fitted with the thermometer must be air-tight.)

The cooling causes the steam to condense. As no air can enter the flask, the water vapour in the flask is at a pressure less than that of the atmosphere. Under this condition the water boils at a temperature below 100° C.

The boiling-point of water is 100° C. at 760 mm. pressure. If the pressure is increased the boiling-point is raised. If the pressure is decreased the boiling-point is lowered. Climbers on high mountains, where the atmospheric pressure is low, find that water boils at a temperature too low to cook some foods properly. The water, although boiling, is not hot enough to cook the food. This was found to be the case on the Everest expeditions.

We have here a method of finding the height of a mountain. The temperature is taken at which water boils and from this the atmospheric pressure is known. We have seen (p. 45) that an estimate of the height of the mountain can then be made.

THE EXPANSION OF GASES

Unlike solids and liquids, all gases expand equal amounts for the same temperature rise; but when a gas is heated it can only expand if there is room for it to do so. If it is in a closed space it cannot expand, so the pressure increases. If we wish to find out how a gas expands on heating, we must arrange for the volumes of the gases to be measured at the same pressure.

Expt. 275. To find the relation between the temperature and volume of a gas, the pressure remaining constant. (Fig. 233.) A is a bulb containing air. The liquid in the tube B is concentrated sulphuric acid. This keeps the air dry. By using the tap C the levels of the acid in both arms of D can be kept the same. The whole is surrounded by a mixture of ice and water contained in the inverted bell-jar. A stirrer S and a thermometer T are in the bell-jar. E is a tube through which steam can be blown into the liquid in the bell-jar.
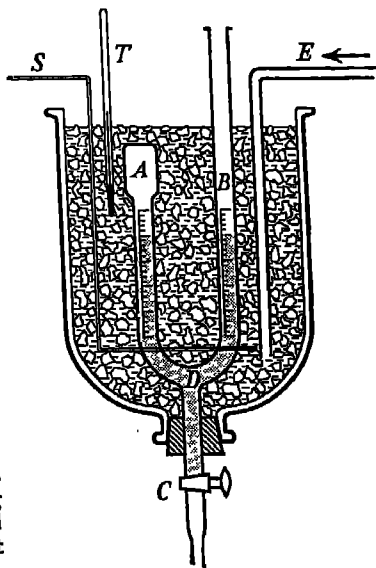
FIG. 233.

See that the columns of acid in both arms of B are level. The air in A is then at atmospheric pressure. Read the volume of the air from the graduation marks on B. This will be the volume of the air at $0°$ C. and at atmospheric pressure. Now blow steam through E. After the ice has melted the temperature rises and the air in A expands. Stop passing the steam and stir well. Adjust the levels in B by means of the tap C so that the pressure in A is again atmospheric. Read the new volume and the new temperature. Pass more steam and repeat the experiment so that you have readings for different temperatures to $100°$ C. Try to get readings for about every $10°$ rise.

Draw a graph of your results. It should be a straight line showing that the increase in volume is uniform.

By this method it can be shown that air expands 1/273 of its volume at 0° C. for every degree Centigrade rise in temperature. This is approximately true of all gases, and the fact is known as Charles's Law.

Consider a volume of gas 273 c.c. at 0° C. When the temperature falls 1° C. the volume will be 272 c.c. (at the same pressure). At −10° C. the volume will be 263 c.c. and so on. Theoretically, at −273° C. the volume will be 0 c.c. Before this temperature is reached, however, the gas will change into a liquid. The temperature of −273° C. is called the *absolute zero* and the temperature scale which has this for its zero is called the *absolute scale* of temperature. The temperature of a body expressed in this scale is called the *absolute temperature*.

CHARLES's LAW may now be expressed in another form: *When the pressure is constant, the volume of a gas varies directly as its absolute temperature.* Expressed mathematically,

$$V \propto T,$$

where V is the volume in c.c. and T is the absolute temperature.

Absolute temperature is, of course, obtained by adding 273 to the temperature measured in the Centigrade scale.

Thus $T = C + 273.$

APPLICATION OF CHARLES'S AND BOYLE'S LAWS

In our study (p. 406) of the behaviour of a gas as the pressure changes, we learnt that the volume of a gas varies inversely as the pressure, the temperature remaining constant (Boyle's Law). That is,

$$V \propto \frac{1}{P} \ (T \text{ being constant}).$$

We now know also (Charles's Law)

$$V \propto T \ (P \text{ being constant}).$$

Therefore $V \propto \dfrac{T}{P}.$

That is, $V = R.\dfrac{T}{P} \text{(where } R \text{ is a constant)}.$

That is,
$$\frac{P.V}{T} = \text{a constant.}$$

That is,
$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

where $P_1 V_1 T_1$ represent one set of conditions and $P_2 V_2 T_2$ represent another set of conditions.

This is known as the gas equation. It is no more than a mathematical expression of the facts of Boyle's and Charles's Laws. For the purpose of calculation the gas equation may be used, but there is no need to use it. It is often better to remember the two laws and apply them directly.

When the volume of a gas is stated, it is necessary to state also the temperature and the pressure. Usually the volume of gas is given as what it would be if the temperature were 0° C. and the pressure 760 mm. This is called *standard temperature and pressure* (S.T.P.), or sometimes *normal temperature and pressure* (N.T.P.). Gases, however, are measured at all temperatures and pressures for the sake of convenience, but the volume at S.T.P. is easily calculated if the volume at any other temperature and pressure is known. An example will make this clear.

Example 1. What is the volume of a gas at S.T.P. which measures 500 c.c. at 15° C. and 740 mm. pressure?

The absolute temperature is 273+15 degrees, that is, 288°.

Method 1. Apply the gas equation:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2},$$

$$\frac{760 \times V_1}{273} = \frac{740 \times 500}{288}$$

where $V_1 = $ vol. at S.T.P.

Therefore
$$V_1 = \frac{500 \times 273 \times 740}{288 \times 760} \text{ c.c.} = 461 \text{ c.c.}$$

Method 2. Apply the laws directly.

The volume of the gas at 0° C. must be less than at 15° C. as the gas contracts on cooling. Therefore the volume of the gas at 0° C. (that is, 273° Abs.) will be

$$\frac{500 \times 273}{288} \text{ c.c., the pressure remaining constant.}$$

The volume of the gas will be less than this at 760 mm. pressure because we are measuring the volume at a greater pressure than 740 mm.

The volume at S.T.P. will therefore be

$$500 \times \frac{273}{288} \times \frac{740}{760} \text{ c.c.} = 461 \text{ c.c.}$$

Example 2. A mass of gas measures 1,000 c.c. at S.T.P. What will it measure at 10° C. and 780 mm. pressure?

Again this may be done by applying the gas equation but we shall do it by the second method.

Consider the pressure. The new volume is at a higher pressure, therefore it will be less than 1,000 c.c. if the pressure alone is considered.

Consider the temperature. The new volume is at a higher temperature, therefore the volume will be greater than 1,000 c.c. if the temperature alone is considered.

Considering both temperature and pressure together, the volume of the gas at 780 mm. pressure and at 10° C. will be

$$1,000 \times \frac{760}{780} \times \frac{283}{273} \text{ c.c.} = 1,010 \text{ c.c.}$$

GAS THERMOMETERS

As gases expand much more than liquids for the same temperature rise, gases should be good substances to use in thermometers. Some gases, e.g. hydrogen, oxygen, and nitrogen, expand uniformly and this is another advantage of using a gas. Gas thermometers, however, are bulky and are not easily carried about, and they are very inconvenient in

general use. They are used in the laboratory for standardizing liquid thermometers.

Gas can be used in a thermometer in two ways. The pressure can be kept constant, then the increase in volume can be used for determining the temperature; or the volume can be kept constant, and the differ-ence in pressure used for deter-mining the temperature. The first method is employed in the Constant Pressure Gas Thermo-meter, and the second in the Constant Volume Gas Thermo-meter.

1. The constant pressure gas thermometer

Fig. 234. *A* is a bulb contain-ing the gas and also a little concentrated sulphuric acid to keep the gas dry. *B* and *C* are two glass tubes connected by rubber tubing *D*. Mercury is poured into *C* and it rises in *B*. By manipulating the tap *E* the levels of the columns of mercury

FIG. 234.

in *B* and *C* can be made the same. When *A* is placed in the liquid, the temperature of which is required, the level in *B* falls because the gas increases in volume, and the level in *C* rises. The tube *C* is then raised until the surfaces of the columns *B* and *C* are level. The gas in *A* is then at the original pressure. The increase in volume has caused the level in *B* to fall, and as *B* is graduated the temperature may be calculated.
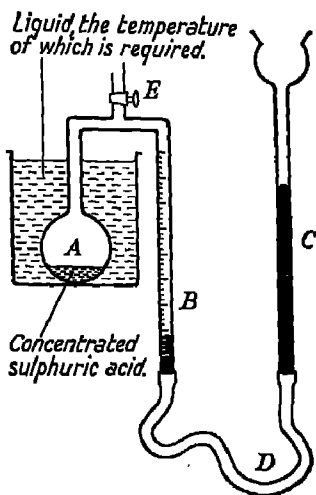
2. The constant volume gas thermometer

The same apparatus may be used as for the Constant Volume Gas Thermometer. When the gas expands and the level in *B* falls, the tube *C* is raised until the level in *B* reaches its former position. The gas is then at its original volume, but

the pressure is increased. This time the tube C is graduated and the temperature may be calculated from the increase in pressure.

CHANGE IN VOLUME OF A GAS WITHOUT THE ADDITION OR SUBTRACTION OF HEAT FROM OUTSIDE

When a gas expands without having any heat given to it from outside, it becomes cooler.

Expt. 276. Allow the air from a bicycle tyre to escape through the valve and put your finger against the issuing jet. Your finger feels cold. The air in the tyre is at the same temperature as that outside if it has been left in the same place long enough, but as it expands on coming into the atmosphere it cools.

When a gas expands work has to be done to push back the surrounding atmosphere and the energy for it is taken from the gas itself in the form of heat. Hence the gas becomes colder. If sufficient heat is supplied from outside the gas, the latter expands without a fall in temperature because the energy required is supplied from outside.

When a gas contracts without having any heat taken from it to the outside the opposite effect is noticed, the gas becomes warmer. The energy previously taken from the gas in the form of heat is now restored as the atmospheric pressure does work on the gas. Hence the temperature of the gas rises.

This change of temperature without the addition or subtraction of heat from outside is very common in the atmosphere and has a great influence on the weather (p. 509).

The heating of the air which comes about merely by a change in volume without the addition of heat from outside is called *adiabatic heating*. When air rises, from whatever cause, it expands because there is a decrease in pressure with height. It must therefore lose heat. For every 100 metres rise in *dry* air there is a fall in temperature of 1° C., and for every 100 metres sinking there is a rise in temperature of 1° C. This is called the *dry adiabatic lapse rate*. It is very important in the study of weather and we shall return to it when we have studied water in the atmosphere.

QUESTIONS

1. Show that the heat from an electric radiator has been originally derived from the sun.

2. Compare carefully, giving details, the expansion on heating of solids, liquids, and gases.

3. Show how the fact that solids expand on heating affect the construction of (a) railway lines, (b) bridges, (c) the piston and cylinder of an internal-combustion engine, (d) hot-water pipes, (e) clocks.

4. Give details of two applications, either in the home or in industry, of the expansion and contraction of substances when heated and cooled.

5. Describe a simple experiment you could perform to show that the amounts of expansion of water and mercury differ when these substances are raised through the same range of temperature.

6. What does a thermometer measure? What is meant by 'Blood heat', sometimes found printed on household thermometers? Compare the Centigrade and Fahrenheit thermometers. Convert (a) 98·4° F. to degrees Centigrade, and (b) $-5°$ C. to degrees Fahrenheit.

7. What is (i) the freezing-point and (ii) the boiling-point of distilled water at atmospheric pressure? How are they affected by the addition of some common salt to the water? What is the effect on the boiling-point of reducing the pressure? What is the temperature of the steam (at atmospheric pressure) coming from boiling salt water?

8. Describe one form of Maximum and Minimum Thermometer and explain how it works. What are its uses?

9. Describe one form of Thermostat used either in the home, factory, farm, or the laboratory.

10. Give a careful description of what happens as water cools from 10° C. to below 0° C. How do these facts affect the freezing of a pond or lake?

11. State Charles's Law and Boyle's Law and give an equation which embodies both. Describe how the truth of one of these laws can be confirmed by experiment.

12. A gas at 15° C. and 750 mm. pressure occupies 100 c.c. What would be its volume at Normal Temperature and Pressure?

13. (a) Explain the following facts: (1) When air issues from the valve of a bicycle tyre the air feels cold. (2) When air is pumped into a bicycle tyre the pump feels hot.

(b) What is adiabatic heating? What is its importance in the study of weather?

XXV

SPECIFIC HEAT AND LATENT HEAT

THE MEASUREMENT OF HEAT

IT is convenient to have a unit of heat by which to measure the different amounts of heat with which we have to deal. We have a 'yard' to measure distance. The 'yard' is the unit of length in the English System, the 'metre' is the unit of length in the Metric System (C.G.S. System).

The unit for measuring heat in the Metric System is called the *calorie*. This is defined as: The amount of heat which will raise the temperature of 1 gm. of water through 1 degree Centigrade.

Thus to raise 1 gm. of water from 0° C. to 100° C., 100 calories are required. To raise 500 gm. of water through 1° C., 500 calories are required. To raise a litre of water from 0° C. to 100° C. would require $(1,000 \times 100) = 100,000$ calories, because a litre of water being 1,000 c.c. weighs 1,000 gm.

A *great calorie* is 1,000 calories.

How many calories will raise the temperature of a gm. of water through b degrees Centigrade? Through how many degrees Centigrade would a calories raise the temperature of b gm. of water?

The unit of heat still used by British engineers is called the *British Thermal Unit* (B.Th.U.). This is defined as the quantity of heat which will raise the temperature of 1 lb. of water 1 degree Fahrenheit. A *therm* is 100,000 British Thermal Units. Gas companies supply gas at a certain cost per therm. When a cubic foot of gas, measured at a standard temperature and pressure, is burned, a certain amount of heat is produced. This varies according to the quality of the gas. Usually it is about 450 B.Th.U. per cu. ft. Gas companies are compelled by law to state how many therms are produced when a definite volume of gas is burned. This is called the *calorific value* of the gas. The companies are not allowed to charge for the gas by volume, but must charge by the therm.

THERMAL CAPACITY

Expt. 277. Immerse a carbon lamp or an ordinary 60-watt lamp in a quantity of water and note the rise of temperature in two minutes. Repeat with the same weight of paraffin oil. The rise of temperature of the paraffin oil is nearly twice as much as that of water.

To raise the temperature of a given mass of paraffin oil 1° C. requires less heat than is required to raise the temperature of the same mass of water 1° C. The amount of heat required to raise the temperature of a body 1° C. is called its *thermal capacity*. The thermal capacity of paraffin oil is less than that of an equal mass of water. Thermal capacity depends on the nature of the substance and its mass.

SPECIFIC HEAT

The specific heat of a substance is the number of calories needed to raise the temperature of 1 gm. of it by 1° C. This is a convenient definition and is suitable for all practical purposes. It is, however, not strictly accurate. The specific heat of a substance is correctly defined as the ratio between the thermal capacity of the substance and the thermal capacity of the same mass of water. Throughout this book, however, we shall use the first definition.

The specific heat of mercury is 0·03 calories per gm. This means that it requires 0·03 calories to raise the temperature of 1 gm. of mercury through 1° C.

The specific heat of water is greater than that of almost any other substance. Usually it is found that the specific heat of a liquid is greater than that of a solid. If the specific heat of a substance is known, then the amount of heat necessary to raise the temperature of a given mass of it a certain number of degrees Centigrade can easily be calculated. For example: The specific heat of lead is 0·031 calories per gm. To raise the temperature of 100 gm. of lead from 16° C. to 26° C. requires 100 × 10 × 0·031 calories, that is, 31 calories. If the same mass of lead cools from 26° C. to 16° C., the same amount of heat, namely, 31 calories, is given out.

TO FIND THE SPECIFIC HEAT OF COPPER

Expt. 278. Obtain a thin aluminium can. The amount of heat taken by such a vessel can be neglected for the purpose of this experiment. Stand it on corks inside another larger can as shown in Fig. 235. Put in the inner can (which may be called a calorimeter) some water at the ordinary temperature of the room. Note the volume of water and its temperature. Tie a piece of previously weighed copper to a thread and suspend it in boiling water for ten minutes. The copper will then be at the temperature of the boiling water throughout its mass. The next stage of the experiment must be performed quickly. Have a thermometer standing in the water in the calorimeter. Take out the copper from the boiling water, quickly shake it to free it from as much water as possible, and put it in the calorimeter. Stir with the thermometer (carefully, or you will break it) and note the rise in temperature. Write down the highest temperature reached. The temperature of the water in the calorimeter should be taken just before the copper is taken from the boiling water.

Fig. 235. A simple calorimeter.

Labels: Aluminium can. Outer can. Air Water Cork

We now have all the facts necessary to enable us to find the specific heat of the copper. The method is called the *method of mixtures*. We work on the assumption that all the heat lost by the copper goes to warm up the water in the calorimeter. That is:

Heat lost by copper = Heat gained by the water.

This is not quite true in the method we have used. Some of the heat lost by the copper goes to heat the calorimeter itself. If, however, we use a very thin calorimeter and stand it on a non-conductor there will not be much heat lost. Some small amount of heat is lost to the air during the time of the transference of the copper to the calorimeter. Again, if we work quickly, this will not be much. The method, however, cannot be relied upon to give a very accurate result, though good enough for most purposes. There are more accurate methods but they are not so simple.

Mass of water .   .   .   . x gm.
Mass of copper   .   .   .   . y gm.
Temperature of water in calorimeter .   .   . $t°$ C.
Temperature of copper   .   .   .   .   . $100°$ C.
Temperature of water after mixing with the copper   . $T°$ C.
Rise in temperature of the water   .   .   . $(T-t)°$ C.
Fall in temperature of the copper   .   .   . $(100-T)°$ C.
Heat gained by the water .   .   .   . $x \times (T-t)$ calories
Heat lost by the copper (Specific Heat $= s$).   . $ys(100-T)$
Therefore $ys(100-T) = x(T-t)$

$$\text{Specific Heat of copper} = s = \frac{x(T-t)}{y(100-T)}.$$

TO FIND THE SPECIFIC HEAT OF PARAFFIN

Expt. 279. The method of mixtures may be used. Use the same type of calorimeter and let it contain a known mass of paraffin. This may be done by weighing the calorimeter first empty and then again when it contains the paraffin. Heat a known volume of water. Place a thermometer in the paraffin. Note the temperature of the hot water and pour it into the paraffin. Stir with the thermometer and note the highest temperature reached. The calculation is, of course, on the same principle as in Experiment 278. That is, the heat gained by the paraffin is equal to the heat lost by the water.

In employing this method of mixtures to find the specific heat of a liquid, the liquid must not combine with, or dissolve in, water. When this happens there is a heat change due to the chemical action and then it cannot be assumed that the heat gained by the liquid is equal to that lost by the water.

WATER EQUIVALENT

Sometimes in our experiments a thick calorimeter is used and the amount of heat given to it cannot be neglected. Even when a thin calorimeter is used, if a very accurate result is required the heat given to it must be taken into account. In such cases it is often convenient to find the *water equivalent*, that is, the weight of water which has the same thermal capacity as the calorimeter. Suppose the thermal capacity of a body (i.e. the amount of heat required to raise the temperature of the body $1°$ C.) is 50 calories. Then the water equivalent of the body is 50 gm. This may be found by experiment, but if the mass of the

body and the specific heat of the material of which it is made are known, it is easily calculated.

Let the Specific Heat of the copper of a calorimeter be 0·095 cal./gm.

Let the weight of the calorimeter be 250 gm.

Then the amount of heat required to raise its temperature 1° C. is 250 × 0·095 calories.

That is, the thermal capacity of the calorimeter is 23·75 calories.

Therefore the Water Equivalent of the calorimeter is 23·75 gm.

LATENT HEAT

Expt. 280. Gently heat some water in a beaker. Take the temperature every minute until it boils and for a few minutes afterwards.

When the water is boiling the temperature remains at 100° C. all the time, in spite of the fact that heat is being given to the water. What becomes of the heat that is supplied to the water while it is boiling? It is clear that heat may be given to a body without raising its temperature. The heat is used in changing the water into steam. The molecules of water are close together when the water is in the liquid state, but much farther apart when the water is in the form of steam. The heat supplies the energy required to force the molecules farther apart.

When water changes its state from liquid to vapour its temperature remains the same but it absorbs heat. This is true of all liquids.

Expt. 281. Put some crushed ice into a beaker and place a thermometer in it. Warm gently over a bunsen flame, stirring all the time. Note that as long as some ice is present, the temperature remains constant at 0° C.

In changing its state from solid (ice) to liquid, water remains at the same temperature but absorbs heat. This is true of all solids which melt.

Heat which is absorbed by a body without raising its temperature is called *latent heat* (latent = hidden). The *latent*

heat of fusion is the heat absorbed by 1 gm. of a body as it changes from a solid to a liquid at the same temperature. The *latent heat of vaporization* is the heat absorbed by 1 gm. of a body as it changes from a liquid to a vapour at the same temperature.

When a gas changes to a liquid, the latent heat of vaporization is given out again, and when the liquid solidifies, the latent heat of fusion is given out.

TO FIND THE LATENT HEAT OF STEAM

Expt. 282. Use Marshall's thick copper calorimeter. This consists of a thick heavy copper or brass vessel surrounded by a felt jacket to prevent loss of heat (Fig. 236). The calorimeter is fitted with an india-rubber bung through which pass two tubes, one wide and central, the other narrower and turned at right angles away from the centre. The central tube is arranged to fit loosely into the bung of a tin vessel in which water may be boiled. Weigh the calorimeter (without the felt jacket). Replace the jacket. Place a thermometer in the calorimeter for a few minutes and take the temperature.

FIG. 236. To find the latent heat of steam.

Boil some water in the tin vessel and when steam is issuing from the bung invert the calorimeter over the tin vessel as shown in the figure. When steam issues from the calorimeter in a strong steady stream, remove the calorimeter. When it is cool measure the volume of the water condensed.

The steam going into the calorimeter is at 100° C. Some of this steam gives up its heat to the calorimeter and raises its temperature to 100° C. In doing this some of the steam is condensed to water at 100° C. The heat that raised the temperature of the calorimeter must be the latent heat given up by the water as it changes from steam at 100° C. to water at that

temperature. Hence the heat gained by the calorimeter must
be equal to the latent heat of the mass of steam condensed.

Let the latent heat of steam be . . . . L cal. per gm.
Temperature of calorimeter, cold . . . . $t°$ C.
Temperature of calorimeter, hot . . . . 100° C.
Volume of water condensed . . . . . x c.c.
Weight of water condensed . . . . . x gm.
Weight of calorimeter . . . . . . y gm.
Sp. Ht. of copper, of which the calorimeter is made . s cal. per gm.
Heat gained by calorimeter . . . $(100-t) \times s \times y$ cal.
Heat given up by steam . . . . . $x \times L$ cal.
Therefore　　$L \times x = (100-t) \times s \times y$.

$$\text{Latent heat of steam} = \frac{(100-t) \times s \times y}{x} \text{cal. per gm.}$$

THE LATENT HEAT OF FUSION OF ICE

Expt. 283. Weigh a calorimeter empty and then half-full of warm
water. Take the temperature. Add ice (previously dried with blot-
ting-paper) until the temperature of the water is about as much below
the temperature of the room as the warm water was above it. Take
the temperature as it begins to rise. Weigh the calorimeter and its
contents.

The warm water lost heat. Some of this was used to melt
the ice and some to raise the water produced by the melted ice
from 0° C. to the final temperature. Study the last paragraph
and then try to calculate the latent heat of fusion of ice. The
true value is about 80 calories.

SOME EFFECTS OF LATENT HEAT

It is often warmer after snow has fallen. As the water in the
atmosphere changes to solid snow the latent heat is given out.
Winter rain also makes the air warmer because the latent heat
of vaporization is released; but summer rain usually makes the
air cooler. This is because the rain is quickly evaporated and
heat is absorbed in the process. It is true that the heat absorbed
in the process is no more than the heat given out as the vapour
turned into rain, but the heat is absorbed from the lower
atmosphere where its loss can be noticed, while the heat given

out was partly lost in the upper regions where the rain was formed.

Water evaporating from the leaves of plants in summer keeps the plants cool. The heat is absorbed largely from the leaves.

Gardeners do not water their plants in the garden during brilliant sunshine because the water quickly evaporates and some of the heat necessary for this purpose is taken from the
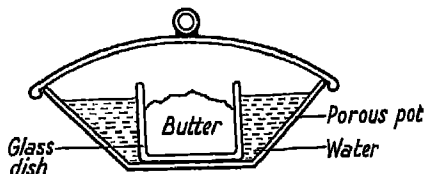
FIG. 237. A butter-cooler.

plants and the ground. Under such conditions some of the more delicate plants may be chilled and their growth checked even if they are not permanently injured.

In summer, butter may be kept cool by keeping it in a butter-cooler (Fig. 237). This consists of a porous pot containing water. The latter gradually passes through the porous pot and evaporates from the sides. The heat necessary is drawn largely from the pot. In this way the water is kept cool. The butter-dish is kept in the water. In the tropics water is cooled by keeping it in a canvas bag. This acts like the butter-cooler.

On leaving water after a bathe in the sea, one often feels colder out of the water than in it. This is more noticeable when there is a breeze. The reason is that the water on the body evaporates, and as it does so it takes heat from the body. We shall see later (p. 497) how the body is cooled by the evaporation of sweat.

EVAPORATION CAUSES COOLING

In the examples given above we have seen that when a liquid evaporates the heat needed is often drawn from the object with which the liquid is in contact. As a result the object

is cooled. This fact may be further illustrated by some simple experiments.

Expt. 284. Pour some methylated spirit or ether on the hand. Wave the hand about so that the liquid quickly evaporates. The hand feels very cold. Any volatile liquid may be used for this purpose. The latent heat of vaporization comes from the hand which, therefore, feels colder.

Expt. 285. Put a thermometer in some methylated spirit and note the temperature. Remove the thermometer and wave it in the air. Again note the temperature. Why has there been a fall?

Expt. 286. Suspend two similar thermometers side by side. To the bulb of one fasten a piece of muslin which has been previously dipped in water at room temperature. Blow on both bulbs with a cycle pump. The thermometer with the wet bulb shows a considerable fall in temperature while there is no fall in the temperature of the dry-bulb thermometer. Try to explain this.

Expt. 287. Put a little ether in a small beaker and stand it on a wet wooden block. Using a pair of bellows, blow air through the ether. This evaporates very quickly, takes heat from its surroundings, and the water on which it stands freezes, and the beaker sticks to the block. (You must not have a flame near the ether and you should, as far as possible, avoid breathing the vapour.)

LIQUEFACTION OF GASES

We know that steam can be liquefied (i.e. condensed) by cooling, and we have seen that sulphur dioxide can be liquefied at the ordinary temperature simply by compressing it. The liquid sulphur dioxide is commonly stored in siphons of strong glass. Carbon dioxide can also be liquefied at the ordinary temperature by pressure, and so can ammonia gas. The gas nitrogen peroxide can be liquefied at atmospheric pressure simply by passing it into a vessel immersed in a mixture of ice and salt.

We see then that both temperature and pressure are used to liquefy gases. It has been found, however, that no matter what pressure is used a gas will not liquefy unless it is below a certain temperature. This temperature, called the *critical temperature*, is different for each gas. The *critical pressure* is the pressure which will just liquefy a gas which is at its critical temperature. The critical temperature of carbon dioxide, for example, is 31° C. This means, if carbon dioxide is above this temperature

no amount of pressure will liquefy it. In this condition it is called a true gas. When the temperature is below 31° C. carbon dioxide can be liquefied by pressure, so that below this temperature carbon dioxide is not a true gas but a vapour. The pressure necessary to liquefy carbon dioxide when it is at its critical temperature, 31° C., is 73 atmospheres. This is the critical pressure of carbon dioxide.

The critical temperature for water is 365° C. Above this temperature it cannot be liquefied by pressure. It is a true gas. Below this temperature it can be liquefied by pressure and is a vapour.

All gases have now been liquefied. By boiling them under reduced pressure very low temperatures have been reached. Air begins to liquefy at —184° C. at 760 mm. of mercury pressure, hydrogen at —253° C., and helium at —269° C. This is not very far from absolute zero. The absolute zero has not yet been reached, but it is claimed that temperatures within 1/1,000 of a degree of absolute zero have been attained by a magnetic method.

LIQUID AIR

Liquid air, having such a low temperature, gains heat very quickly if kept in ordinary vessels and soon evaporates completely. It is therefore kept in a special flask known as a *dewar* flask constructed on the principle of the vacuum flask (p. 473). Loss of heat is thus prevented as far as practicable.

When liquid air is allowed to boil, the nitrogen, being more volatile than the oxygen, evaporates first and the remaining liquid gradually becomes richer in oxygen. Advantage is taken of this fact to manufacture oxygen and nitrogen on a commercial scale by the method known as *fractional evaporation*. It can be done very cheaply, and by modern methods of liquefaction and evaporation practically pure oxygen and nitrogen can be obtained.

Practically pure nitrogen comes off first. As the boiling continues some oxygen comes off and so the escaping gas becomes gradually richer in oxygen. At last the gas escaping is practically pure oxygen. By taking these gases in fractions any

mixture can be obtained from almost pure nitrogen to pure oxygen.

REFRIGERATION

We have seen that food can be preserved for a long time by keeping it at a very low temperature. This is essential in hot
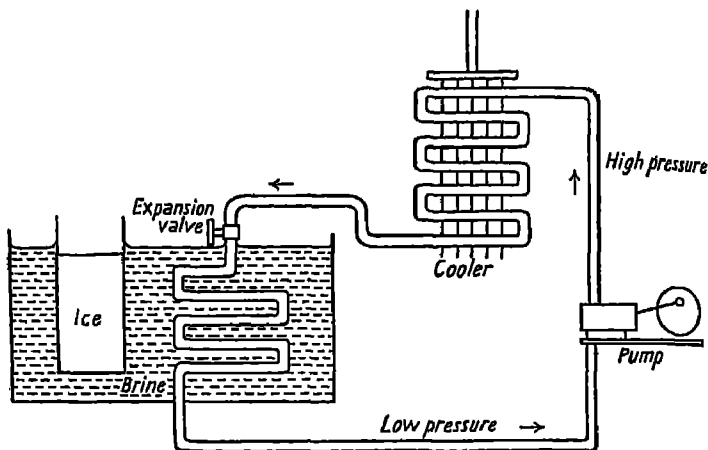
Fig. 238. The principle of a refrigerator.

climates. One method is to keep the food on ice or in a box surrounded with ice which is regularly renewed. This has been largely superseded by refrigerators.

The substance generally used in refrigerators is ammonia which needs only a pressure of 10 atmospheres at ordinary temperatures to liquefy it. Carbon dioxide is sometimes used but this needs a greater pressure. The principle employed is that the ammonia liquid is allowed to evaporate under reduced pressure and the heat necessary for this evaporation is taken from the surroundings which consequently become colder.

Fig. 238 illustrates the main features in one kind of refrigerator. Ammonia circulates in a pipe which has two coils, one of which is in a cooler while the other is immersed in a tank containing brine (or sometimes calcium chloride solution). In the circuit there is an expansion valve where the

coil enters the brine tank and a pump situated between the tank coil and the cooler coil. When the pump is working the expansion valve operates so that the ammonia in the coil passing through the cooler is at a higher pressure than that in the rest of the circuit. Under these conditions it is liquid. The work done by the pump in this part of the circuit appears as heat and this is taken away by the cooler. When the ammonia passes through the expansion valve it enters that part of the circuit which is kept under reduced pressure by the action of the pump. The ammonia evaporates, taking the heat necessary for this change of state from the surrounding brine which is thus cooled. Cans of water are kept in the solution and when the water is frozen the ice is removed. The solution can be made to circulate round a room which is thus kept cold, and may be used for cold storage. The solution can also be made to circulate so that it keeps water frozen to make an ice-rink. The process is continuous, for the same ammonia is compressed again in the condenser tubes. No ammonia is lost.

The modern household refrigerator depends on the same principle, namely, that evaporating ammonia takes latent heat from the surroundings, but modifications have been introduced to avoid having moving parts, thus doing away with noise.

QUESTIONS

1. Define and explain (a) Calorie, (b) Great Calorie, (c) British Thermal Unit, (d) Therm. What is meant by 'Calorific value' of a sample of a gas?

2. Explain the statement: 'The specific heat of lead is 0·031 calories per gram.' Give a brief outline of an experiment you could perform to find the specific heat of copper.

3. A pan containing 2 kg. of water at 15° C. is put on a gas-ring and the water boils in 5 minutes. The water is then thrown away and replaced by 2 kg. of ice at 0° C. The pan is heated on the same gas-ring. How long will it be before the ice is melted and the water formed all boiled away? Neglect the heat given to the pan. (Latent heat of fusion of ice = 80 cal./gm. Latent heat of vaporization of water = 536 cal./gm.)

4. A copper ball of 10 gm. weight is placed in a fire for a long time and then quickly transferred to 100 gm. of water at 15° C. contained in an aluminium calorimeter (the water equivalent of which may be neglected). The resulting temperature of the water is 20° C. What was the temperature of the fire? (Sp. Ht. of copper = 0·1 cal./gm.)

5. Account for the fact that winter rain often makes the air warmer while summer rain often cools the air.

6. What is the difference between a gas and a vapour? How are gases liquefied?

7. Describe one kind of refrigerator. Briefly describe how it works and state the physical principles on which it depends.

8. Define the following: Latent Heat of Fusion, Latent Heat of Vaporization, and give illustrations of their use in the home.

XXVI

TRANSFERENCE OF HEAT

CONDUCTION

HOLD a piece of wire in a bunsen flame. In time the end held in the hand feels hot. The heat travels from the flame along the wire to the hand. We say that the metal of the wire has conducted the heat to the hand, or, the heat has travelled by *conduction*. Use a glass rod instead of a metal wire. It is a long time before the hand feels any heat. The glass is not so good a conductor of heat as the metal. Glass is a bad conductor of heat. Most metals are good conductors, silver and copper being very good, while lead and mercury are not

FIG. 239. Water is a bad conductor of heat.

so good. Even lead and mercury are very good compared with such substances as glass and water.

Expt. 288. To show that water is a bad conductor of heat. Put a little ice in a test-tube and add water until the test-tube is about three-quarters full. Keep the ice at the bottom by means of a piece of metal gauze. Heat the water at the top by a bunsen flame (Fig. 239). In time the water will boil at the top but the ice will remain unmelted at the bottom. Of course, in time the ice will melt because water does conduct heat to some extent and the ice will also gain some heat from the room.

Expt. 289. To show that some metals conduct heat better than others. Arrange three wires of different metals but of the same length and thickness on a piece of wood (Fig. 240). Pin down over them a sheet of 'heat sensitive paper'. Heat the wires at *A* with a bunsen flame and note the colour change as the heat is conducted from the flame along the wires. It will be found that the heat travels along some of the wires farther than along others. The experiment illustrates roughly the fact that different metals have different heat conductivities.

SOME APPLICATIONS OF THE CONDUCTION OF HEAT

Expt. 290. Gradually lower a piece of wire gauze on to the flame of a bunsen burner (Fig. 241). The flame stays below the gauze. Hold another cold gauze about 2 inches above an unlighted bunsen
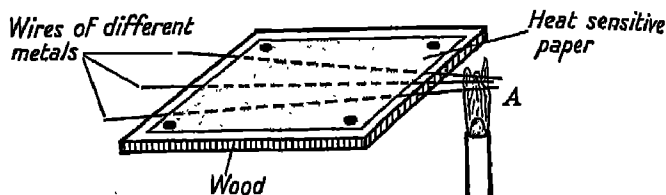
FIG. 240.
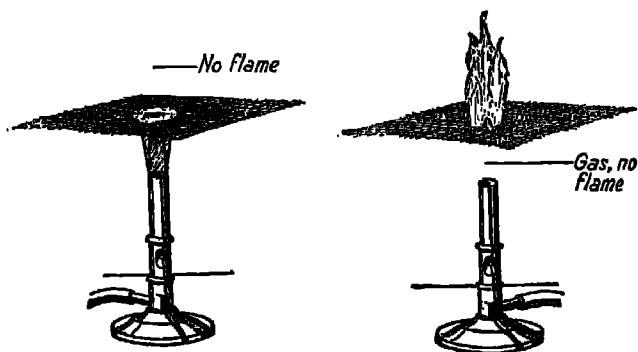
FIG. 241.

and turn on the gas. Light the gas above the gauze. The flame does not strike down below the gauze. Now lift up the gauze. What is the result?

The explanation is that the metal of the gauze is a good conductor of heat. It conducts the heat away so that the temperature above the gauze, in the first experiment, never reaches the temperature of ignition of the gas. Similarly, in the second experiment, the temperature below the gauze is too low to ignite the gas.

This fact is used in the construction of the miner's safety lamp, called the *Davy lamp* (Fig. 242). The gas in most coal-mines is inflammable. A naked light would cause an explosion. Sir HUMPHRY DAVY invented a safety lamp in which the flame from the oil lamp is enclosed in a gauze cage. The tempera-ture outside the lamp never reaches the point at which the outside gas would explode. If for any reason the gauze be-comes hot there is danger and the lamp must be put out. This safety lamp has made possible the working of mines which previously were unsafe owing to the amount of inflammable gas in them. Sometimes the methane (the inflammable gas in a coal-mine) travels through the gauze and burns inside the lamp with a pale blue flame.

Support for top of lamp

Cylindrical wire gauze

glass cylinder

Oil reservoir

FIG. 242. A miner's safety lamp.

When the miner sees this flame he knows that there is more inflammable gas in the mine than usual and that precautions, such as extra ventilation, are neces-sary.

If a piece of metal in the open air is touched on a cold day it feels cold. This is because the metal is much colder than the hand and takes heat from it very quickly, being a good con-ductor. If wood is touched in the same circumstances it does not feel anything like so cold as the iron. It is really at the same temperature as the iron, but it is a poor conductor and does not take heat away from the hand so quickly. In very frosty weather a moistened finger will freeze on to a piece of iron but not on to a piece of wood.

Pans and kettles must be made of good conductors so that the heat can get to the contents: but if the handles are also made of the same material they become too hot to use. The handles are therefore made of some bad conductor such as

wood or horn. A soldering-iron has a wooden handle, a bad conductor, but the end is copper because it is a good conductor.

Clothing must be made of bad conductors. Wool is a bad conductor because air is entangled among the fibres. Air itself is a very bad conductor. Two layers of thin material on the body are 'warmer' than one thick layer of the same material, because of the layer of air between. Wool is 'warmer' than cotton because more air is entangled between the fibres.
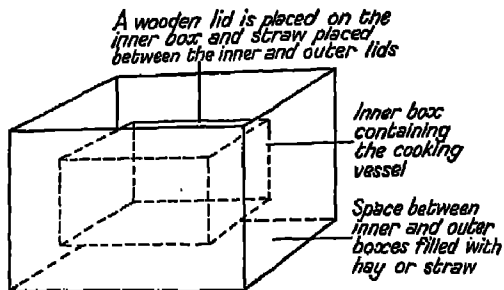
FIG. 243. A hay-box cooker.

'Cotton-wool', however, is very 'warm' because much air is enclosed. It feels 'warmer' getting into bed between blankets than between sheets, because the cotton of the sheets is a better conductor than the wool of the blankets. The fibres of wool are more irregular than those of cotton and so are able to enclose more air.

Asbestos is a very useful substance because not only is it a very bad conductor of heat, but it will not burn. It is used to cover hot-water pipes to prevent loss of heat.

In frosty weather straw is sometimes placed on football fields to protect the ground from frost. The straw is a bad conductor and prevents loss of heat from the ground to the cold air.

Fig. 243 illustrates the construction of a 'hay-box' cooker. The food is partially cooked on a fire or stove and is placed in the cooker which is then closed. It is thus surrounded with non-conducting material (hay or straw) and the heat is kept in

the box and the cooking is completed. Food can be kept warm in such a box for a long time, and there is some saving of fuel.

A THEORY OF CONDUCTION

How is it that heat is able to travel through bodies? The particles of the body do not move carrying the heat with them. The following explanation is generally accepted.

Heat is a form of energy. When a substance becomes hotter, the molecules, which are always moving to and fro (vibrating) vibrate much more vigorously. Consider a metal rod heated at one end. The molecules at this end vibrate much more quickly, more collisions occur between the molecules, and this increased vibration is communicated to the neighbouring molecules. These in turn cause their neighbours to vibrate more quickly. In this way the increased vibration of the molecules is passed along the rod. It must be made clear that the molecules do not travel along the rod, they merely communicate their increased vibration to the other molecules.
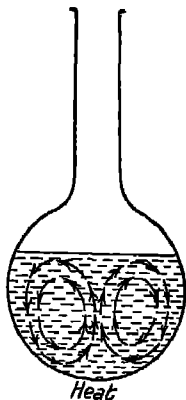
Fig. 244. Convection currents in water.

CONVECTION IN LIQUIDS

Expt. 291. To some water in a flask add a few crystals of potassium permanganate. These colour the water, and, if the latter is not otherwise disturbed, currents in the water may be easily noted when they occur.

Heat the water gently with the flame of a bunsen (Fig. 244). Watch the water as the heating continues. The coloured water can be seen to rise from the bottom, its place being taken by water from the top which flows down. Draw a diagram of the currents in the water. Instead of potassium permanganate you may use fine sawdust or finely divided carbon.

The currents thus set up in the water are called *convection currents*. The water at the bottom expands on heating and thus becomes less dense, or 'lighter'. It therefore rises, while the colder, and therefore denser, water at the top sinks. In convection

the heat travels through the liquid by means of the moving particles of the liquid. Unlike conduction, the particles of the liquid move from one end to the other. It is not a question of the vibration of molecules, but of the movement of particles. Convection in a solid is impossible because the particles are not free to move about as they are in a liquid.

When water is heated in a test-tube it is heated at the bottom. Owing to convection, the whole is soon hot. When in Experiment 288 we heated the water at the top, convection currents could not be formed and the whole of the water took a long time to get hot.

The water in a pond is heated by the sun from the top of the pond. This causes the top layer to become hotter and therefore less dense. This warm layer therefore will not sink. The further heating of the water in a pond must depend largely on conduction, and as water is a bad conductor of heat, it will be a very slow process. The water of a still pond which has become very cold during winter takes a long time in the summer to become warm. The deeper the pond the longer it will take. The cooling of a pond to 4° C. is helped by convection currents, because the upper layers become cooler and denser and sink.

Coloured water

Heat

FIG. 245. Circulation of water by means of convection currents.

APPLICATIONS OF CONVECTION CURRENTS IN WATER

Expt. 292. Fit up an apparatus as shown in Fig. 245. Put some coloured ink in the water in the top vessel. Heat the flask, gently at first. Convection currents are formed and the water rises up the bent tube to the top vessel. The coloured water from the top flows down the vertical tube to take the place of the rising water. Soon all the water in the apparatus is coloured.

It is on this principle that the hot-water supply system of a house works. Examine Fig. 246 and follow the circulation of the water from the boiler behind the kitchen fire. Note the
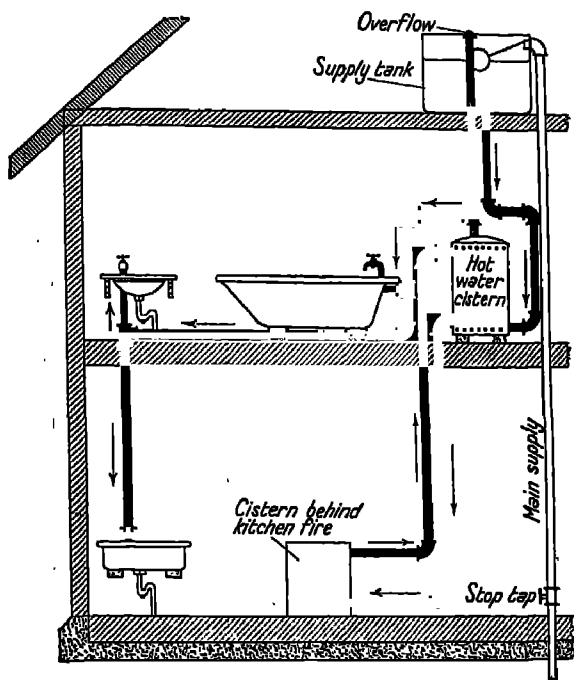
FIG. 246. Hot-water supply system of a house.

arrangement for supplying fresh cold water when some water has been drawn from a tap. Note also the arrangement to prevent an explosion if the water boils and steam is formed.

The heating of a building by the circulation of hot water is also on the same principle. Examine Fig. 247. The water is heated in a boiler in the basement. The hot water rises to the different stories of the building, and after circulating in the pipes round all the rooms, returns, much cooled, to the boiler.
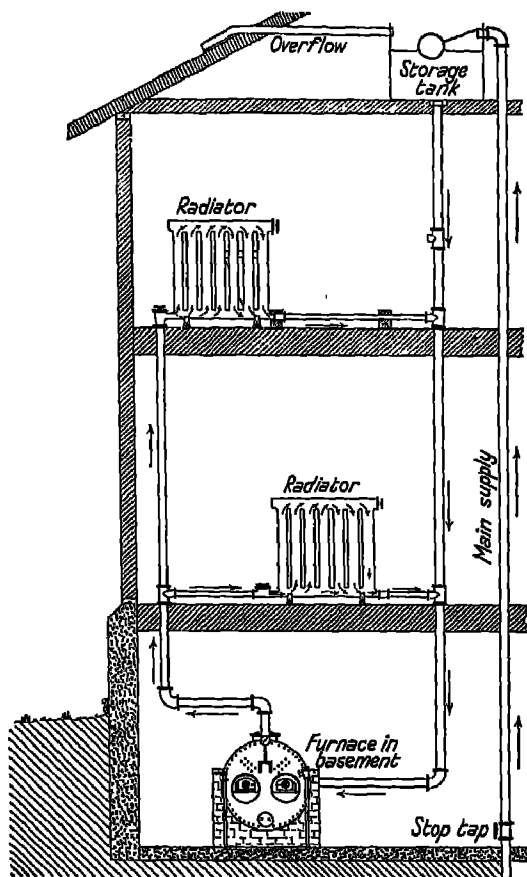
FIG. 247. Central heating by convection currents in water.

The pipes in the boiler-house are usually covered with felt or asbestos to prevent loss of heat before the water reaches the parts of the building to be warmed. In the rooms, the water-pipes are, of course, uncovered so that as much heat as possible may be given to the rooms. To increase the amount of heat given to a room the water is made to pass through 'radiators' so that a

greater surface of the hot metal is exposed to the room. How 'radiators' work will be left to a later paragraph (p. 471). There is often an arrangement in each room, usually against the 'radiator' by which the water may be prevented from going through the 'radiator' when the room is not in use.

This is not the only method of using water to heat a building, but it was once very common and is still in use.

FIG. 248. Convection currents in air.

CONVECTION IN GASES

The particles of a gas, like those of a liquid, but unlike those of a solid, can move about freely amongst themselves. For this reason heat can be transferred in gases by convection currents. This may be illustrated as follows:

Expt. 293. Place a lighted candle at the bottom of a tall lamp chimney standing on a bench. After a time the light goes out. The heat of the flame causes the air to rise, and owing to the upward current in a narrow chimney other air cannot come down to take its place. Thus the burning candle is deprived of oxygen for combustion and the light is extinguished.

Now divide the lamp chimney into two parts by means of a piece of cardboard or thin wood (Fig. 248). Relight the candle. It continues to burn indefinitely. Hold a piece of burning brown paper over the top of the chimney. The smoke from the paper travels down one part of the chimney and up the other, showing the direction of the convection current.

Formerly, coal-mines were ventilated on this principle. There were two shafts to the mine. At the bottom of one a fire was lighted. This caused the air to rise up the shaft. Air travelled down the other shaft to replace that which came up the first, and was directed round the mine before reaching the up-draught shaft. Expt. 294 illustrates the method.
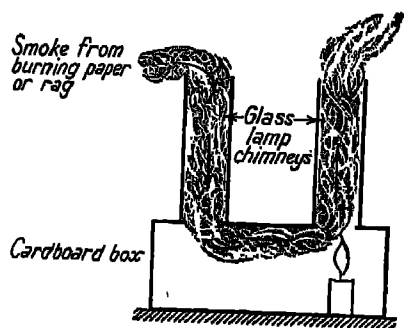
FIG. 249. Convection currents in air.
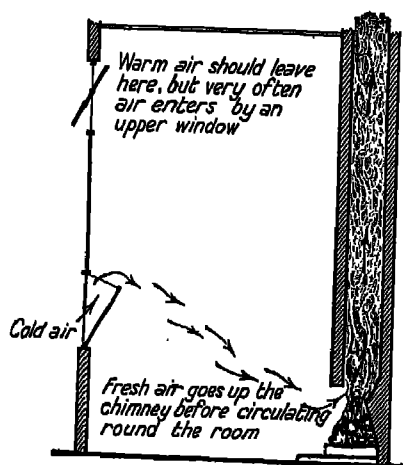
FIG. 250. A poor but very common method
of ventilation.

Expt. 294. Arrange a box and two lamp chimneys as shown in
Fig. 249. Light a candle under one chimney. Hold burning brown
paper above the other chimney. In time smoke will be seen travelling
up the first chimney.

The ventilation of a house depends to some extent on convec-

tion currents. The subject is bound up with the heating of the house. If a single coal-fire is used, it causes air to rise up the chimney and air is drawn into the room to replace that gone up the chimney. For this reason a seat in front of the fire is sometimes the draughtiest seat in the room. Where 'radiators'
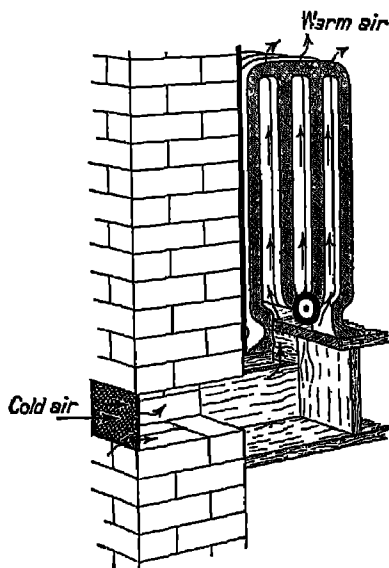
FIG. 251. Cold fresh air warmed as it enters a room.

are used, convection currents are formed against each 'radiator'. Fresh cold air is drawn into the room, and it is often arranged that this should come through a cavity behind the radiator and so be warmed as it enters. Where there is no heating of the room, convection currents are still formed. The air given out by human beings is usually warmer than that in the room. It therefore rises and escapes through the window or ventilator. This assumes that there is no breeze blowing towards the window. Windows open at the top are intended to allow the warm and partially used air to escape. Often, however, they allow fresh air to enter.

Convection currents in the atmosphere play a great part in producing weather conditions. The hot air near the equator rises and there is a tendency for colder air to flow in from the north and south. This, however, is modified by other factors, one of which is the rotation of the earth on its axis. Air over sea and land may be unequally heated and so cause convection currents. Winds are convection currents on a large scale, but heat is not the only factor in determining the direction of the wind. Mountains and the spin of the earth on its axis have to be taken into account. See also Chap. XXX.

How the Heat of the Sun reaches the Earth

How does the heat of the sun reach the earth? Between the earth and the sun there is for the most part 'empty space'. The air does not reach more than a few miles high and then there is nothing for about 93,000,000 miles before the sun is reached. Conduction of heat is impossible, for there is nothing to conduct it. Convection also must be impossible for there are no particles of matter to form convection currents. Heat is emitted by the sun and is absorbed by the earth, yet there is no material substance between the sun and the earth. The heat is said to reach the earth from the sun by *radiation*. The method is explained by the *wave theory*. It is assumed that pervading all space is the *ether*. This is not a material substance and has nothing to do with the chemical substance ether. It has no weight and no chemical properties. It is not matter in the ordinary sense of the word. The ether is supposed to exist between the molecules of all matter. The molecules of a solid are very small and are separated from each other by considerable distances compared with their size. The molecules of a liquid are much farther apart, and those of a gas still farther apart. The molecules vibrate in the ether, and it is this vibration which we know as heat.

Expt. 295. Fill a pneumatic trough with water and float a cork on the surface near the side. Wait until the surface of the water is calm, then plunge a stick in the water as far as possible from the cork. Continue with gentle plunging. Waves will form in the water and soon they will have travelled to the cork which will be set in motion by the wave.

WAVE-MOTION

We must pay some attention to this subject now as it is so important in the radiation of energy, and we shall need to understand it not only in our study of heat but in light, sound, and electricity. Wave-motion is conveniently studied by means of a ripple tank. This is done in Chap. XXXII.

Consider first the waves in water. If corks are put in a straight line on the surface of some water in a large trough, the
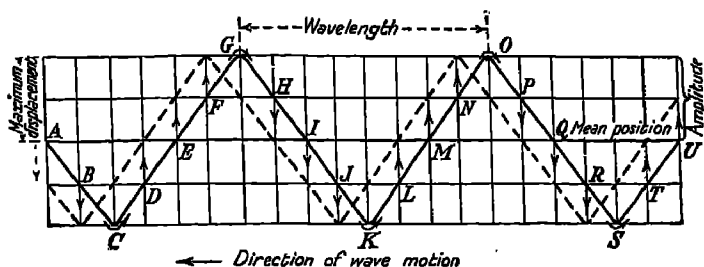
FIG. 252. Transverse wave motion.

motion of the corks can be observed when waves are caused to travel across the water. The only motion observed is the up and down motion of the corks. They do not move horizontally, and yet something *seems* to move horizontally across the surface of the water. When the cork is at its highest position it is said to be on the crest of a wave, when at its lowest position it is in the trough of a wave. If the water could be suddenly frozen at a particular moment, the positions of the corks would be as in Fig. 252. *A* is in the position it would occupy if the whole surface of the water were at rest. *B* is a little lower down, while *C* is in the trough. *D* is at the same height as *B*, and *E* as *A*, but *D* and *E* are moving upwards while *A* and *B* are moving downwards. *C* has just finished moving downwards and is about to move upwards. *F* is higher than *E* and *G* is at the crest. *F* is moving upwards and *G* has finished moving upwards and is about to move downwards. *I* is just as high as *A* and is moving in the same direction. *J* is like *B* in these respects.

The *displacement* of a cork is its distance from the mean position, that is, the position it would occupy if the water were at rest. The maximum displacement is from the mean position to the crest (or trough). The cork *A* moves from the mean position to the trough and then from the trough through the mean position again to the crest. It then moves downwards to the mean position again. In doing this it makes one *vibration* or *oscillation*. It continues vibrating in this way and so do all the other corks, but each cork is a little in front of the one behind it in the line.

Consider the corks *A* and *I*. Their displacements are the same, their speeds are the same, and they are travelling in the same direction. They are said to be in the same *phase*. *B* and *J* are also in the same phase, and so are *C* and *K* (in troughs) and *G* and *O* (on crests). The distance between any nearest two corks in the same phase is called the *wave-length*. We are using corks in our illustration for convenience, but they serve only to indicate the position of particles. We can say, then, that the wave-length is the distance between any two nearest particles in the same phase.

Now consider what happens when each cork is displaced one stage farther, that is, *A* goes to the present height of *B*, and *B* goes to the present height of *C*, and so on all along the line. A broken line has been drawn in the figure through the new positions of the corks and an ordinary line through their original positions. It is seen that the crest of the wave has moved to the left. The only motion of the corks has been at right angles to *AQ* but the wave has travelled in the direction of *QA*.

The time of one vibration, that is, the time taken by each individual particle to complete its path, is called the *periodic time*, and the number of complete vibrations made by a particle in one second is called the *frequency*. The maximum displacement of each particle from its mean position is called the *amplitude*. (See Pendulum, p. 324.)

Hence the periodic time (in secs.) = 1 sec./Frequency.

The *velocity* of a wave is the distance travelled by the crest

(or any part of the wave) divided by the time taken to travel that distance.

$$\text{Velocity} = \frac{\text{Wave-length}}{\text{Periodic time}}$$

because the crest (say) travels one wave-length while a particle is making one vibration.

Hence,   Velocity = Wave-length × Frequency.

We have spoken only of waves in water, but all this applies to wave-motion in the ether. In the waves we have considered, the motion of the particles is always at right angles to the direction in which the wave is moving. This is the case in all wave-motion in the ether. It is called *transverse* wave-motion. When we study sound, we shall have to consider another kind of wave-motion, but this is not in the ether but in a material medium.

The sun radiates energy into space and this travels as wave-motion in the ether. Some of these waves have a very great wave-length, nearly one million centimetres. Others have very tiny wave-lengths, some as small as 0·0000000007 cm. The wave-lengths of light are between 0·00004 cm. and 0·00008 cm. All energy, however, is transmitted through the ether at the same velocity no matter what its wave-length. This is the velocity of Light. When we study this subject we shall have to return to the consideration of wave-lengths.

ABSORPTION AND RADIATION OF HEAT

Heat which reaches us from a radiating body (*radiant heat*) does so without warming the intervening medium very much. In this respect radiation is different from conduction and convection. The air absorbs a little of the heat radiated from the sun and so becomes warmer, but the heat that reaches the ground has not heated the air.

When a body emits heat, it radiates it in all directions. Any matter coming in the path of the rays will absorb some of the heat. It is found that dull surfaces are better absorbers than bright surfaces and that dull surfaces are better radiators than bright surfaces. Good absorbers are good radiators. 'Radiators'

in buildings are often painted a dull colour, that is, a non-shiny colour. Dull white is almost as good a radiator as dull black. The heat from a 'radiator', however, comes into the air much more by convection than by radiation, so that it does not matter much how they are painted.

It is interesting to consider the greenhouse from the point of view of radiation. Experiment shows that the amount of radiant heat transmitted by glass differs according to the temperature of the radiating body. Most of the heat radiated from a white-hot body passes through glass, but not much of the heat radiated from a red-hot body will pass through. Hardly any from a body at a temperature below 100° C. will be transmitted. The heat from the sun passes into a greenhouse through the glass. The heat is absorbed by the plants and soil, the temperature of which rises. These objects now radiate heat, but little or none of this passes through the glass because the temperature of the radiating bodies is less than 100° C. In this way a greenhouse becomes hot. There is also another factor in the warming of a greenhouse which has no connexion with radiation. As the greenhouse is closed it is sheltered from the convection currents outside.

THE TRANSFERENCE OF HEAT

We have seen that heat can be transferred in three ways: Conduction, Convection, and Radiation. In conduction there is no actual movement of the particles of the substance. In convection there is movement of the particles. Conduction and convection act through material media, but radiation is through the ether.

Usually heat is transferred by all three methods together, but sometimes one method predominates. In the heating of the water in a pond most of the heat reaches the surface of the pond by radiation from the sun, but some of it is by conduction from the warm air above it. The heat travels through the water by convection and conduction (p. 461).

The air in a room heated by hot-water pipes receives some of its heat by radiation from the pipes, some by convection

currents set up in the air by contact with the hot pipes, while conduction occurs between the metal of the pipes in contact with the hot water.

In preventing loss of heat from a body it is necessary to take precautions against loss by all three methods. This is well illustrated in the construction of the vacuum flask. Fig. 253 shows a vacuum flask in section. The flask is double walled, the
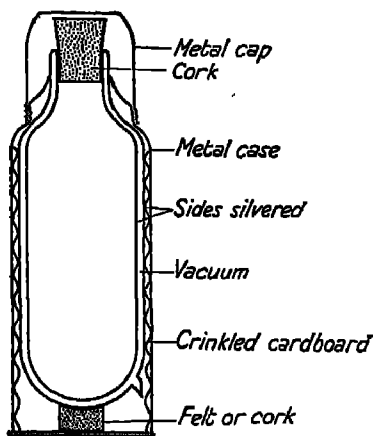
FIG. 253. A vacuum (or thermos) flask.

space between being exhausted of air. This prevents loss of heat from the inside by convection currents, and also by conduction through the air. Air is a very bad conductor of heat, but some conduction would take place if the space between contained air. The flask is also made of glass which is a bad conductor. The two inside surfaces facing the vacuum are highly silvered. Radiant heat will pass through a vacuum but it is reflected back from the highly polished surface. A polished surface also radiates less heat than a dull one. As both walls are silvered this reflection works both ways. The packing of the flask in the container is done in such a way that loss of heat by conduction is reduced to a minimum. The flask stands on cork which is a bad conductor, and is packed round the sides (to prevent

breakage by vibration) with non-conducting material such as cork or crinkled paper.

A flask so made will not only keep a hot liquid warm for a long time, it will keep a cold liquid cold. This is because the precautions taken to stop heat from travelling out also act the other way and will not let heat travel in. The precautions tend to prevent the passage of heat either way. As there are no perfect non-conductors, in time the contents of the flask will be at the same temperature as the air surrounding the flask. If great care is taken in the construction of the flask it will keep things hot or cold for a very long time.

The walls of many houses are built with an air cavity between them, that is, there is a space between the outer and inner walls. This serves as a heat insulator, keeping the house cooler in summer and warmer in winter. (The circulation of air through the cavity also tends to keep the house dry.)

THE COOLING SYSTEM OF A MOTOR-CAR

As another example of the transference of heat let us study the cooling system of a motor-car engine (Fig. 254). In doing its work the petrol engine develops a great amount of heat and unless much of this heat is lost by the engine very quickly, it will become so hot that it will not work efficiently. The cooling system is designed to help the engine to lose its heat. The engine proper is enclosed in a water jacket which is part of a closed system through which the water may flow. Included in the system is the radiator. When the engine is working, heat is lost by conduction through the metal casing to the water in the jacket. The water becomes hotter and therefore less dense. It therefore rises and a circulation through the system is effected, colder water taking the place of the water in the jacket. In its path the hot water flows through the radiator. This is designed to allow the water to lose as much heat as possible during the time it is flowing through. For this reason its passage through the radiator is long and the pipes are fitted with flanges to increase the surface in contact with the air. The water loses heat by conduction to the metal of the radiator and this loses it mostly by conduction to the surround-

ing air. A little is also lost by radiation. 'Radiator' is a bad name as the majority of the heat is lost from it by conduction. The amount of air passing through the fins is increased by a fan worked by the engine. This draws air through the fins. The work of the fan is more important when the car is stationary than when it is running, for then there is a current of air
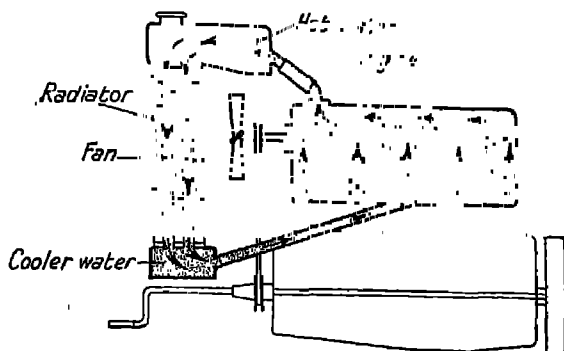
FIG. 254. Cooling system of a motor-car engine.

passing through the fins due to the motion of the car through the air.

Thus the cooling of the car is effected by employing all three methods of transference of heat: conduction from the engine to water and from the metal of the radiator to the air, convection currents in water, and some radiation from the metal of the radiator to the air.

Formerly it was difficult to keep the petrol engine of a car cool enough, but the engines have been improved so that now there is seldom any trouble through over-heating. Indeed sometimes the cooling system is so good that it is difficult to keep the engine warm enough, and on very cold days part of the radiator is usually kept covered. (See p. 487 for the Cooling System of an Aero-engine.)

During severe frost there is some danger that the water in the cooling system will freeze if the car is kept in an unheated

garage. If this happens the radiator may be damaged by the ice, because water expands on freezing. To prevent freezing the cooling system may be emptied of water by using a tap provided for this purpose. Sometimes glycerine or some other substance is added to the water in the cooling system which prevents the water from freezing by lowering the freezing-point.

OCEAN CURRENTS

The currents in the ocean are due to many and various causes. Some are due to convection currents in the water caused by the unequal heating of the surface. It was formerly thought that this was the chief cause, but it is now realized that most of the great ocean currents are due to the winds. The currents are generally in the direction of the prevailing winds and when these change according to the season the currents change with them. There are also some currents due to differences in the salinity of the water and this causes difference in density. Evaporation in very hot areas increases the surface salinity, and rain and great rivers decrease the salinity in other areas. Nevertheless convection currents do play a part. The cold, and therefore the dense, water at the poles tends to flow in an undercurrent to the equator, and the warm, and therefore less dense, water at the equator tends to flow in a surface current towards the poles. The currents are deflected by the land masses and also by the rotation of the earth. Only the surface of the water is directly heated, and water being a bad conductor, the lower depths of water are not heated except very slowly. These convection currents are thus very weak and are easily masked by the stronger currents due to the winds and other causes, but that they do occur is proved by the fact that the deeper water at the equator is always very cold.

LAND- AND SEA-BREEZES

The specific heat of water is greater than that of land. Thus when the sea and land are subjected to equal amounts of radiation from the sun, the earth becomes warmer more quickly than the sea. On the other hand, when both sea and

land are cooling, the land cools more quickly than the sea. Consider what happens in calm weather when the sun is shining in a clear sky over an area consisting of land and water. When the sun rises, the land is cooler than the sea because it has been cooling more quickly than the sea during the night. Both
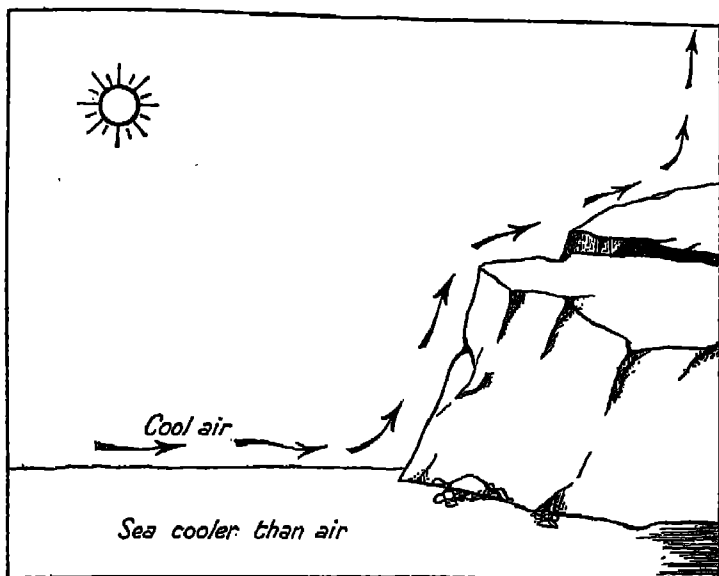
FIG. 255. A sea-breeze.

the land and sea become warmer, and soon the land is at the same temperature as the sea because it becomes warmer more quickly than the sea. As the sun continues to shine, the land becomes warmer than the sea, and the air above the land becomes warmer than that above the sea. This occurs partly by conduction from the land to the air, partly by convection, and partly by long-wave radiation from the earth to the lower regions of the air. Under these conditions the air above the land rises and an upper current flows from the land to the sea, while a lower current of cooler air flows in from the sea

(Fig. 255). This lower current is felt on the land, and is known as a sea-breeze.

As the sun approaches its setting position both the land and the sea lose more heat than they receive from the sun, and thus begin to cool. This continues through the night. The sea,
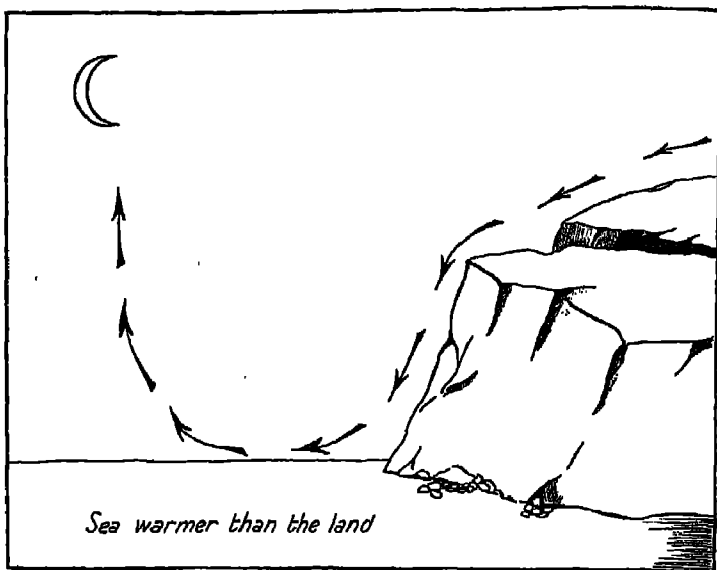
FIG. 256. A land-breeze.

however, cools less quickly than the land, and soon the air above the sea is warmer than that above the land. An upper current of air now flows from the sea while a lower current flows from the land over the sea. This is known as a land-breeze.

In some parts of the world, land- and sea-breezes occur with great regularity. They are more likely to occur where there is a great range of temperature between day and night. If the sky is covered with cloud, land- and sea-breezes will not be produced. This is because in the daytime a great deal of the sun's radiation is absorbed by the cloud before it reaches the earth and at night the clouds act as reflectors for the radiation

from the land and so help to keep its temperature higher than would otherwise be the case.

Land- and sea-breezes are not very strong and are due to local causes; hence the existence of strong winds, or winds over a large area due to other than local causes, will mask the breezes. Land- and sea-breezes are often noticed in England during periods of calm settled weather. They are a common feature of the climates of tropical islands.

The Temperature at the Top of High Mountains

At the top of a mountain the air is nearly always colder than at the foot. A man ascending in a balloon finds that the temperature decreases as he ascends. This is a well-known fact but one which on first consideration appears strange. It is sometimes asked: why should the air be colder at high altitudes when it is nearer to the sun? But it is not nearer to the sun in any practical sense. Even if we are on a mountain 5 miles high we are only 5 miles nearer to the sun and this can be of no importance when we remember that the sun is about 93,000,000 miles from the earth. Within the limits we can reach on earth, nearness to the sun cannot have any influence on temperature.

Another question often asked is: why is it colder on the top of a mountain considering that the rays of the sun have travelled through so much less air and so much less heat will have been absorbed? It is true that the rays of the sun are more powerful on the top of a mountain than at the foot. People sunburn much more quickly at high altitudes than they do in the plains. But the direct heat of the sun is not the main factor in heating the air. Only a fraction of the radiant heat from the sun is absorbed by the air. When a fire has just been lit in a cold room, its heat can be felt on the face, yet the air in between is still cold. Sometimes when there is a cool breeze on a sunny day, we feel warm when we are in the sun, but we feel cold immediately a cloud obscures the sun. The rays of the sun heat us and the earth, but do not heat to any great extent the intervening air. This fact may be proved by a direct experiment.

Expt. 296. On a calm day suspend a thermometer in the open air so that the direct rays of the sun fall on it.· Note that the temperature rises. Now wave the thermometer about in the air, at the same time allowing the sun's rays to fall on it as before. Note that the temperature falls. In the first case the sun heated the thermometer; in the second case the same happened, but the thermometer lost some of its heat to the air as it was waved about. In the first case the air in the immediate vicinity of the thermometer was heated by the thermometer and then prevented further loss of heat.

As the sun's rays pass through the atmosphere some are absorbed, and to this extent the air is warmed, but although the total amount of heat absorbed is considerable it does not cause a big rise in temperature because there is such a great mass of air.

The air is heated by contact with the earth which is heated by radiation from the sun. Consequently the air nearer to the surface of the earth is warmer than that farther away. Convection currents also play their part in heating the air, but as the source of the heat causing the currents is the earth, the temperature falls as the height increases. When the air in contact with the surface of the earth is heated by conduction from the earth, the air rises because its density decreases. As the air rises the pressure decreases and the air expands. When air expands owing to a reduction of pressure the temperature falls. (See Dry Adiabatic Lapse Rate, p. 442.)

When air containing water vapour rises, the rate of fall of temperature is slowed down by the addition to the air of heat given out from the water vapour as it condenses to water (*Latent heat of vaporization*). We shall deal with this when we study Water in the Atmosphere. The net result is that temperature falls (in the temperate regions) about $0.5°$ C. for 100 metres ascent (or $2.7°$ F. for 1,000 ft.).

QUESTIONS

1. Distinguish carefully between conduction, convection, and radiation. Give one example of each type to be found in the home.

2. Explain the following observations: (i) the flame in a Davy lamp does not cause an explosion in a mine where gas is present; (ii) on a cold day in the open air a piece of iron feels much colder than a nearby piece of wood; (iii) the handle of a metal teapot is usually made of some non-metallic substance.

3. Which will keep you warmer, three layers of thin clothing or one layer of the same material, the thickness of which is equal to the three combined? Give a reason for your answer.

4. If you wish to heat a tank of water by using an electric heating element, where would you put the element, in the middle of the tank or at the bottom? Give reasons for your answer.

5. Draw a labelled diagram of either (a) the hot water system of a house, or (b) the heating of a building by the circulation of hot water, and explain how the system works.

6. Explain how (i) hot-water 'radiators' heat a room, and (ii) the ventilation of a room is affected by an ordinary coal fire.

7. Why is the air in an unheated greenhouse usually warmer than that outside?

8. Give a labelled diagram of a vacuum flask showing its construction, and explain how it prevents loss of heat by conduction, convection, and radiation. Will a vacuum flask also keep liquids cold? Give a reason for your answer.

9. Internal-combustion engines must not be allowed to become too hot. Give details of the cooling system of a motor-car engine. State briefly how the cylinder of a motor-cycle engine is prevented from becoming too hot.

10. What are 'Land and Sea Breezes'? Explain their occurrence.

11. Give a careful explanation of the fact that the temperature at the top of a high mountain is lower than that at its foot.
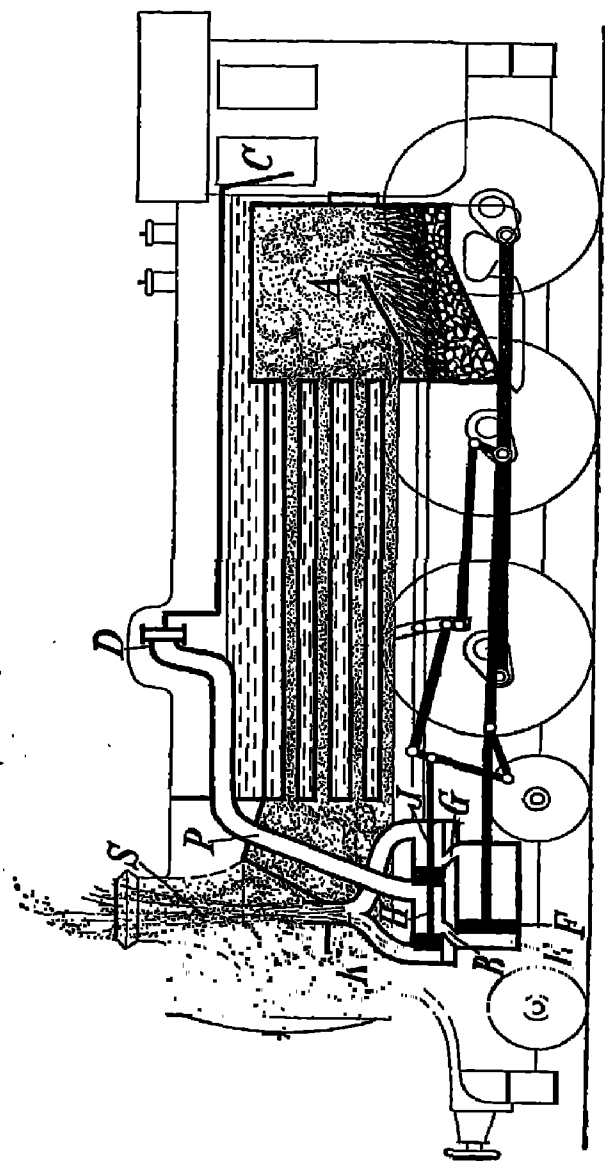
FIG. 257. A railway steam-engine.

XXVII

MECHANICAL EQUIVALENT OF HEAT

HEAT AND MECHANICAL ENERGY

WE have seen that when friction occurs heat is produced. In our study of machines we found that some of the work put into a machine was lost owing to the production of heat. Thus mechanical energy can easily be changed into heat energy. This is because heat is really molecular energy. When we hammer a piece of iron, both the hammer head and the iron become warmer. Some of the kinetic energy of the moving hammer head is transformed into heat energy. The molecules of the iron are now vibrating more quickly, and this increase in the kinetic energy of the molecules of the iron has been obtained from the kinetic energy of the hammer head.

It is often a disadvantage when mechanical energy changes into heat energy, for in this way some of the energy is lost for useful work. It is a disadvantage in machines, and it is prevented as much as possible by reducing the friction of the moving parts to a minimum. To change heat energy into mechanical energy is, however, a useful transformation. Two well-known examples are the steam engine and the internal-combustion engine. In both these engines fuel is burnt to produce heat and this is ultimately converted into mechanical energy. We shall illustrate this process by considering briefly how a steam locomotive and a petrol engine work.

THE STEAM LOCOMOTIVE

Coal is burnt in the fire-box A (Fig. 257) and the hot gases from the fire pass along the flues to the smoke-stack S where they escape through the funnel. In their passage they heat the water in the boiler which is in the form of tubes, thus allowing a great heating area. Since the boiler is closed the steam gradually increases in pressure as the heat increases. In the type we are describing the pressure is allowed to rise to about 15 atmospheres. There is a safety-valve which opens and

allows the escape of steam when this pressure is reached. We often hear the steam escaping from this valve when a locomotive is waiting in a station.

There is a regulator in the dome D, controlled from the cab C. By this means steam is allowed to escape into the steam-pipe P, which leads from the dome, through the smoke-stack to the steam-chest. The steam enters the cylinder E through the passage B behind the piston F. The steam expands, pushing the piston forward. This operates the piston-rod which turns the wheel. As the wheel moves, it operates a valve-rod J connected to a valve H which closes the passage B and at the same time opens another passage G from the steam-chest allowing steam to travel into the cylinder on the other side of the piston, thus forcing it back. At each stroke of the piston, forwards and backwards, the used steam is forced out of the cylinder through an exhaust tube K which leads to the exhaust in the smoke-stack. The exhaust is fitted with a nozzle through which the steam passes with great force. This increases the draught in the funnel and causes the coal to burn more quickly, thus increasing the heat produced. The 'puffs' of the engine are made by the used steam passing out of the exhaust nozzle. There is one puff for each stroke of the piston whether backwards or forwards. The to-and-fro motion of the piston-rod is converted into the circular motion of the wheels by the crank-shaft.

It is clear that all the heat produced by the burning of the coal is not converted into mechanical energy. Heat is lost (1) by hot steam leaving the engine; the latent heat of vaporization is thus lost; (2) by hot gases from the fire leaving the engine through the funnel; (3) by friction of the moving parts; and (4) by radiation from the boiler.

THE PETROL ENGINE

We shall consider the common petrol engine used in a motor-car or a motor-cycle. The energy is obtained by burning petrol. Fig. 258 illustrates a simple 'four-stroke' engine. The petrol passes from the tank to the carburettor where it is vaporized and mixed with air. From there it passes through

the inlet valve I into the cylinder C, the exhaust valve being closed. The piston P travelling down the cylinder draws the petrol vapour into the cylinder. This is called the *inlet stroke*. The piston then travels up the cylinder, both valves being now closed. This compresses the mixture of gas and air into a very small space. This is the *compression stroke*. When the piston
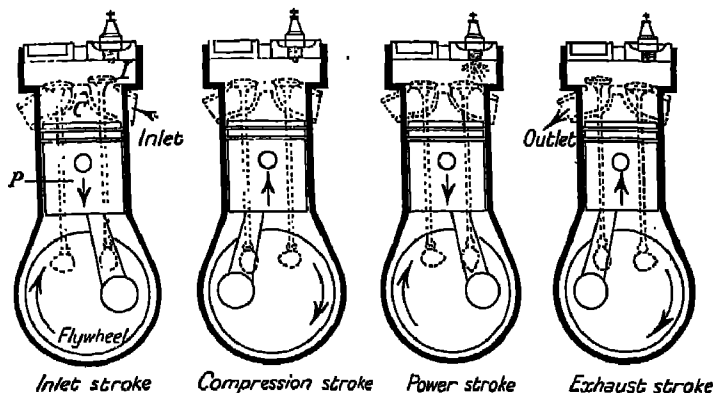
FIG. 258. Diagram of a four-stroke motor-cycle engine.

is at the top of this stroke, an electric spark, caused by the sparking-plug, explodes the compressed mixture. Great heat is generated and the gases expand forcing the piston down the cylinder. The gases formed are carbon dioxide, carbon monoxide, and water vapour. These occupy about the same space as the mixture of petrol vapour and air from which they are made, but the gases are at a very high temperature, and this causes a great expansion. This stroke of the piston is called the *power stroke*. At the end of the power stroke the piston returns up the cylinder, the exhaust valve being open and the inlet valve being closed. This sweeps the used gases out of the cylinder. This is called the *exhaust stroke*. The cycle of operations is then repeated. There is one power stroke for every four strokes of the piston.

The vertical motion of the piston is converted to a circular motion of a flywheel by a crank-shaft which is connected through the gearbox to the back wheel of the motor-cycle.

Much heat is lost in this type of machine. The heat is so great in the cylinder that special arrangements have to be made to carry it away as rapidly as possible, otherwise the piston becomes so hot that it expands too much and excessive friction is caused. It might even stick in the cylinder. We have seen how the cylinders of a motor-car engine are cooled (p. 475). In a motor-cycle the cylinder is cooled by the rush of air past it as the machine travels along. The outside of the cylinder is extended into flat projections which give a greater cooling area to the cylinder.

As in a steam engine, all the heat produced by the burning petrol is not converted into mechanical energy. Some is lost when the hot gases leave the exhaust pipe and more is lost by conduction and radiation from the cylinder. The petrol engine is used to drive many different types of vehicle, especially motor-cars, motor-cycles, and aeroplanes.

AERO-ENGINES

Aeroplane engines differ in detail from motor-cycle and motor-car engines, but they are the same in principle. The main difference is in the number of cylinders. A motor-cycle has one, sometimes two, cylinders and a motor-car may have six or more. An aero-engine has many more. This is because there is a limit to the horse-power which can be obtained from one cylinder. It is also found that smoother running is obtained when a number of cylinders is used.

The cylinders may be arranged *in-line* or *radially*. In the former case the cylinders are arranged one behind the other in the same crank-case and each piston operates a separate crank. If there are more than six cylinders, two (and sometimes four) rows are made. They are all connected to one crank-shaft.

Radial engines have their cylinders arranged around the crank-case. Seven or nine cylinders are on one crank but there may be two or three separate banks of cylinders.

The engines are cooled either by air or a liquid cooler. In air cooling the cylinder is fitted with fins, as on a motor-cycle engine, to aid cooling; baffles and deflectors are also used to direct the air on to the parts of the engine which need it.

In liquid cooling the liquid used may be water alone or a mixture of water and another liquid which prevents freezing in cold weather. This liquid is sometimes used alone without water. The cooling liquid is pumped through a jacket surrounding the cylinders and then through a radiator. The heat from the engine is taken away by the liquid and then dissipated to the air as the liquid flows through the radiator. It is very much like the cooling system of a motor-car engine (p. 475) except that convection currents are sufficient in a motor engine but pumping is necessary in an aero-engine. It is also more elaborate in the aero-engine owing to the increased complexity of the engine.

In both systems of cooling, air or liquid, the air eventually receives the heat. In air cooling it does so directly, but in liquid cooling the liquid acts as an intermediary.

We have seen (p. 363) that the lubrication of an engine is very important. In an aero-engine lubrication reduces friction and wear and tear. It thus prevents the production of more heat by friction and the loss of power due to the moving parts overcoming friction. If the lubricating system fails heat is generated in the parts of the engine which are very difficult to cool and the metal is damaged. Even when the moving parts are properly lubricated much heat is generated. For this reason the oil is supplied under pressure to some parts of the engine and this ensures a quick flow of oil through the parts carrying away the heat. Other parts receive the oil in the form of a spray caused by the oil being splashed about by the moving parts.

JET PROPULSION. THE GAS TURBINE ENGINE

The traditional aero-engine, as we have seen, is a piston engine operating a crank-shaft which turns an airscrew, thus providing the thrust. In the gas turbine engine the thrust may

be produced by the ejection of hot gases at a high velocity from a nozzle near the tail. The details of this machine are beyond the scope of this book but we can study the underlying principles.

No doubt you have often played with a balloon shaped like a sausage. If the air is allowed to escape from one end of
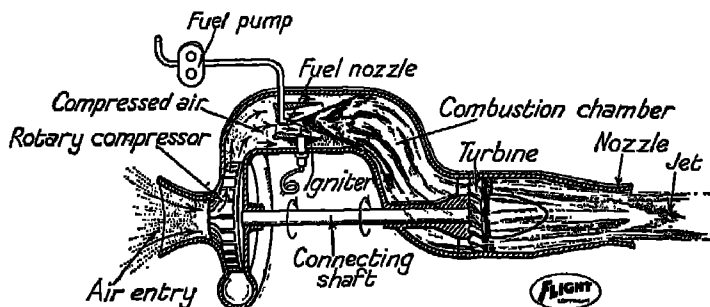
,   Fɪɢ. 259. Cycle of operations in a gas turbine engine.

the balloon, it rushes round until it is deflated. It is propelled by the *reaction* of the jet of air. It is somewhat similar to the recoil of a gun when a projectile is fired (p. 323). The jet of air acts like the gun and the balloon like the projectile. If a man were to stand in a small boat and throw stones over the stern into the water the boat would move forwards. This principle has been known since Newton's time and is usually stated: *Action and reaction are equal and opposite.*

A Roᴄᴋᴇᴛ is operated by creating a continuous jet of gases which, by the reaction, thrusts it in the opposite direction to the jet. Rocket propulsion, however, must be distinguished from jet propulsion. In the former the hot gases are produced by a combustible substance reacting with another substance containing oxygen. Thus the rocket carries with it, not only its fuel, but its oxygen supply in the form of a compound rich in oxygen and easily decomposed. The rocket is thus independent of atmospheric oxygen. So far as combustion is concerned there is no theoretical difficulty in sending a rocket to the

moon. In jet propulsion, the fuel is carried with the engine and the oxygen of the air is used for its combustion. Thus an aircraft propelled by a jet, unlike a rocket, cannot travel outside the atmosphere.

Fig. 259 illustrates very simply the cycle of operations in a gas turbine engine providing jet propulsion.

1. Air is drawn in at the front of the engine and compressed in a chamber. This is done by means of a rotary compressor working on a shaft.

2. Fuel is sprayed into the compressed air from a fine nozzle. The fuel may be petrol, paraffin, heavy oil, other hydrocarbons, or other substances rich in hydrogen and carbon.

3. The fuel burns continuously in this compression chamber. A sparking-plug (called a primary igniter) is used to ignite the mixture at first, but once the engine is started the sparking-plug is no longer needed.

4. The mixture of heated air and the gaseous products of combustion turns a turbine by forcing its way at high velocity through the blades. The turbine and rotary compressor are on the same shaft. Thus the energy of the turbine is used to draw in and compress the air.

5. After passing through the turbine the mixture of hot gases is discharged at high velocity through a nozzle at the back.

Once started the whole cycle of operations is continuous. The engine is started in the same way as a piston engine in a motor-car. An electric motor rotates the shaft, air is drawn in and compressed. The fuel is ignited by a sparking-plug and the engine starts. The sparking-plug is then shut off. The engine continues to work as long as fuel is injected into the chamber. A throttle, regulating the amount of fuel entering the chamber, controls the speed of the engine. It will be noted that only part of the energy produced by the combustion is used for ejecting the gases. A portion is used in driving the turbine which drives the rotary compressor.

A gas turbine engine has many advantages over the piston engine. It is much simpler to construct and therefore cheaper

to manufacture. As there are fewer moving parts it is less liable to mechanical breakdown and not much energy is lost due to friction. Cheaper fuels can be used and there is also less risk of fire with them. A gas turbine engine gives much greater power in relation to weight than a piston engine. Its fuel consumption is higher at speeds below about 500 m.p.h., but above that speed it is just as economical of fuel, while its performance at great heights and at high speeds is much better.

The gas turbine engine is, like the piston engine and the steam engine, a heat engine. It is a device for transforming the chemical energy in the fuel, first to heat energy, and then to mechanical energy.

THE MECHANICAL EQUIVALENT OF HEAT

When heat energy is transformed into mechanical energy, how much of the latter can be obtained from one unit of the former?

We have seen that much heat is wasted when a steam engine is used to transform heat energy into mechanical energy, but if a steam engine could be made which lost no heat, how much heat would be required to give a definite amount of mechanical energy?

We know from our study of the conservation of energy that there must be a definite relation between the number of heat units used and the number of mechanical units obtained. This was not always known, because heat has not always been looked upon as a form of energy. As soon as it was suspected that heat was a form of energy then it was seen that there must be a definite relation between the amount of heat used and the amount of mechanical energy obtained, and vice versa.

The first man to find this relation was JAMES PRESCOTT JOULE (1818–89). He started to work on the problem about the year 1840 and by 1849 he had established that there was a definite relation between the amount of work done and the heat generated. He found that to produce 1 B.Th.U. of heat it was necessary to do 772 ft.-lb. of work. Later, other workers were able to repeat Joule's experiments on a larger scale and

with more accurate thermometers and they found the result to be: 778 ft.-lb. = 1 B.Th.U. This result is now accepted as correct. It is known as the *mechanical equivalent of heat*, or, in honour of the man who first found it, *Joule's equivalent*. It is often denoted by J.

It is not necessary in this book to go into the detail of Joule's method, but some indication of it can be given (Fig. 260).
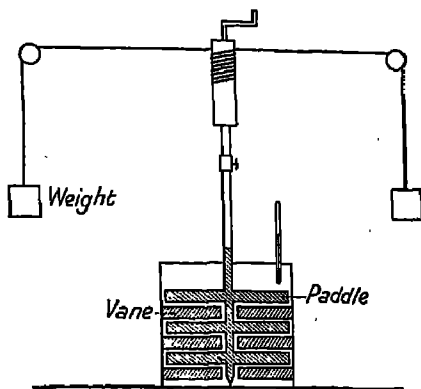
FIG. 260. Joule's calorimeter.

Joule caused paddles to rotate in water in a copper calorimeter which was fitted with fixed vanes so that the water would not swirl round the calorimeter. The friction of the paddles and vanes against the water caused the temperature of the water and the calorimeter to rise. From this, the heat gained by the water and the calorimeter was calculated. The paddles were rotated by a spindle operated by falling weights just heavy enough to rotate the paddles against the frictional resistance of the water. When the weights reached the bottom, he wound them to the top again (without causing the paddles to rotate) and allowed them to fall again. This was continued until the temperature of the water had been raised sufficiently. The work done was calculated in ft.-lb. by multiplying the total weight by the distance fallen and then by the number of times

they fell. All this sounds very simple, but to obtain an accurate result Joule had to take many precautions and apply many corrections.

QUESTIONS

1. Taking a steam locomotive as an example show how heat may be converted into mechanical energy. State in what ways much of the heat produced by the burning of the coal is lost.

2. Describe, with the aid of labelled diagrams, the working of an internal-combustion engine. Is this a heat engine? Give reasons for your answer.

3. What do you understand by the term Mechanical Equivalent of Heat? Briefly describe Joule's method of finding a value for it.

4. In boring a hole through a piece of metal, enough heat was produced to raise the temperature of 10 gm. of water from 50° F. to 105° F. What was the amount of work done?

If the work took two hours to perform, at what h.p. did the drill work?

XXVIII

HEAT AND LIVING THINGS

THE EFFECT OF EXTREMES OF TEMPERATURE ON LIVING THINGS

GREAT heat destroys all forms of life. Very few organisms can exist at a temperature of 50° C. and even those that are adapted for life in hot springs, perish at about 80° C. This temperature is much less than that of boiling water. The bacteria, many of which cause disease, are of microscopic size and these may be destroyed if heated to the temperature of boiling water. If a needle is held for a short time in a flame any micro-organisms on it are destroyed. The needle is sterilized and may safely be used for pricking the finger. Tinned foods keep fresh and wholesome because all the bacteria which cause decay and disease in human beings are killed by heat before the food is tinned. After heating, the tins are sealed and the food thus kept out of contact with the air, and so from fresh infection.

Some foods are sealed in the tin or glass container first and then kept for some time at a temperature that ensures the death of all bacteria. At the present day nearly all sealed food is reliable, but if the tin seems bent inwards or blown out, the contents should be treated with suspicion. It is possible that some bacteriological activity has been going on, causing a chemical action resulting in a change of volume.

Extreme coldness destroys all life. Some organisms, however, can withstand very low temperatures provided that the actual cell contents are not frozen or that ice is not formed in the tissues. Some bacteria which cause decay in meat or other foods can withstand low temperatures for a long time. *Cold storage* of food preserves it, not because it kills the bacteria, but because it reduces their activity. Protoplasm can function only between certain temperatures. At low temperatures bacteria cannot reproduce themselves and life generally is almost at a

standstill. When the food is removed from cold storage the bacteria resume their activity, and decay of the food sets in unless it is quickly used.

Mammoths, a species of hairy elephants that became extinct some 20,000 years ago, have been dug out of the ice and have been found preserved by the cold so that the flesh was still edible.

Freezing turns potatoes black and peas brown and makes them leathery and inedible. In *quick freezing* only small crystals of ice are formed and these do not break the cell walls, so that when the contents thaw the mineral salts and other substances do not escape. Nowadays one can get quick-frozen peas, fish, strawberries, peaches, lobster, spinach, and chicken. Foods have been kept in this way for ten years, but for ordinary commercial purposes probably two or three years should be taken as the maximum time. Vitamins in vegetables are not destroyed during this process.

Generally speaking warmth increases life activities while cold slows them down. For example, seeds do not germinate until they have sufficient warmth. A cold spring means slow germination of the seeds. Plants also grow more quickly when warm, provided the heat is not too great and that they have all the other conditions suitable for growth. With regard to animals it is not always apparent that cold reduces their activities. Man, for instance, is often stimulated into greater activity in the cold. He takes more vigorous exercise in the winter, and on a cold day usually walks faster than on a warm one. This extra activity releases energy in the form of heat, which maintains the protoplasm at the temperature at which it carries out its functions most efficiently.

'COLD-BLOODED' AND 'WARM-BLOODED' ANIMALS

You have probably noticed that you are usually warmer than your surroundings even on a summer's day. This property does not belong to man alone but to all mammals and birds. If you have ever held a living bird in your hand you have probably noticed that it feels warmer than your hand, as

indeed it is. On the other hand, if you pick up a frog or a fish, it feels cold.

Animals are divided into two classes: *warm-blooded animals* and *cold-blooded animals*. The terms, however, do not describe the conditions properly. It is found that the so-called warm-blooded animals have a constant temperature when in health. A healthy man, for example, on hot days or cold days, always has the same temperature. It is about 98° F. (37° C.) This may very easily be tested by putting the bulb of a thermometer in the mouth or under the arm-pit. This temperature is so constant that the doctor looks upon a rise or fall in the temperature as a sign of ill health. Our temperature rises when we have a severe 'cold' and we are then 'feverish'. Warm-blooded animals are more suitably called *constant temperature animals*.

Cold-blooded animals are not always cold. The temperature varies with that of the surroundings. A frog buried in the mud during the winter is very cold, but the same frog leaping about in the summer is much warmer. A fish in a pond in winter is about as cold as the water, while in the summer it is as warm. A lizard in the hot desert is approximately as warm as its surroundings. A better name than cold-blooded is *variable temperature animals*.

ANIMALS

CONSTANT TEMPERATURE. 'Warm-blooded'. Mammals and birds.	VARIABLE TEMPERATURE. 'Cold-blooded'. All the lower animals, e.g. frogs, fish, worms, &c.
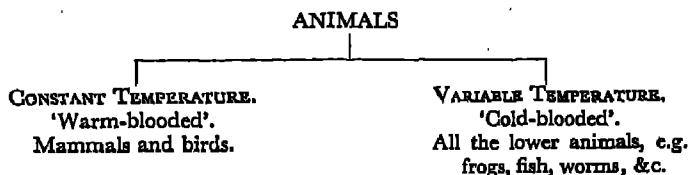
HIBERNATION

Many variable-temperature animals such as the frog stop the search for food in the winter and hibernate. They find a suitable place and apparently go to sleep. They neither eat nor move. Respiration is carried on very slowly, so slowly, in fact, that it is difficult to detect the breathing movements in some animals. The temperature is nearly the same as that of their surroundings, and at this low temperature the speed of all

the chemical processes in the body is slowed down, making it impossible for the animal to lead an active life. Even if an active life were possible at the temperature, food is either scarce or absent altogether so that the animal would die if it did not hibernate.

A few mammals such as the hedgehog, dormouse, bat, and field-mouse hibernate during the winter. In the summer they behave like other 'warm-blooded' animals but as the winter approaches they retire to a sheltered place, their temperature falls, but never so low as that of their surroundings, and they spend the winter in a kind of sleep. The rate of respiration falls, that of the dormouse from 80 per minute in the summer to between 10 and 15 when hibernating. They do not feed, the fuel for respiration being supplied by the fat of their own bodies. This again is an advantage to these animals, as the food they eat is scarce in winter. The squirrel is a partial hibernator because it wakes up occasionally and feeds on food it has stored in preparation for the winter.

The Regulation of Temperature in Man and Other Constant-Temperature Animals

To maintain a constant body temperature it is necessary for an animal to be able to regulate both the production of heat in the body and the loss of heat from the body. The regulation is usually unconscious, but in the case of man it is sometimes conscious. Movement of the muscles produces heat, and in cold weather the body temperature may be kept up by increased exercise of the muscles. Man often 'keeps himself warm' by vigorous exercise in the winter.

Man also consciously prevents loss of heat from the body by wearing clothes. Other animals have some kind of protective layer for this purpose. The bird has feathers, the dog hair. Both these materials entangle air and so make a good nonconducting layer near the skin. Some animals have a layer of fat below the skin. The fat layer in the whale ('blubber') is particularly well developed.

If vigorous muscular exercise is taken during hot weather the body temperature tends to rise. Let us consider how this

is prevented in the case of man. In the first place man may do something consciously to prevent the rise in temperature. He may work more slowly or put on lighter clothing. Even if he takes no action consciously, his temperature will remain constant. This is helped by the sweat glands and the condition of the blood-vessels in the skin. The evaporation of sweat keeps the skin cool because the skin supplies the heat for evaporation. This cools the blood in the capillaries in the skin, and the cooled blood circulates through the body so that the cooling effect of the skin is felt all over the body. When the body is hot the blood capillaries in the skin increase in diameter and, therefore, contain more blood. Remember how the face 'reddens' when you feel hot. In this way heat can be lost through the skin to the air outside by the usual methods of transference of heat. During cold weather, when the body needs to conserve its heat, the blood capillaries of the skin are constricted so that the blood-supply is reduced. This prevents loss of heat from the body. The 'blueness' of the skin during cold weather is due to this cause. Sometimes the skin becomes so cold that the circulation in some of the exposed and extreme parts of the body, such as the fingers and nose, almost stops. It is then that 'frost-bite' may occur.

Some mammals have few or no sweat-glands and are often covered with hair which prevents, to some extent, the action of the skin in cooling the body. The dog has very few sweat-glands. It cools itself by increasing the rate of breathing ('panting') so that the surface of the lungs is cooled. The blood carries the cooling effect through the body. A dog often sits with its tongue out in hot weather. The evaporation of the saliva cools the surface of the tongue and so helps to keep the temperature of the body from rising.

Birds have no sweat-glands. They seek the shade of trees during hot weather, but an increased rate of breathing is probably the chief method of keeping the temperature down. In cold weather the feathers are fluffed out, so entangling more air, which acts as an additional non-conducting layer.

QUESTIONS

1. How and why is heat used in sterilization and in the preservation of some foods? What is the effect of keeping meat in a refrigerator?

2. What do you understand by the terms 'cold-blooded animals' and 'warm-blooded animals'? What better terms could be used for these classes of animals. Give two examples of each class.

3. What is hibernation? Name two mammals and one other animal which hibernate, and describe the conditions of the animals during hibernation. What is the advantage of hibernation to these animals?

4. Describe in detail how the temperature of man is regulated. Compare this with the regulation of temperature in two other mammals. Why does a bird fluff out its feathers in cold weather?

XXIX

WATER IN THE ATMOSPHERE
WATER-VAPOUR IN THE ATMOSPHERE

WE know that there is water-vapour in the atmosphere even on a 'dry' day. Anhydrous copper sulphate turns blue when exposed to the air, and substances like calcium chloride, cupric oxide, and sulphuric acid take moisture from the air.

Expt. 297. Weigh some freshly fused calcium chloride which has been cooled in a desiccator, and then expose it to the air for about an hour. Note that it has taken moisture from the air and has dissolved in it. Reweigh. The difference in weight is the weight of water taken from the air.

Apart from experiment we know that there must be water in the air. Water evaporates from puddles in the road and the vapour must go into the air. We also see water coming out of the air as rain, dew, or snow. Is there a limit to the amount of water-vapour the air can hold? Can the air become saturated with water-vapour? What controls the amount of water-vapour in the air?
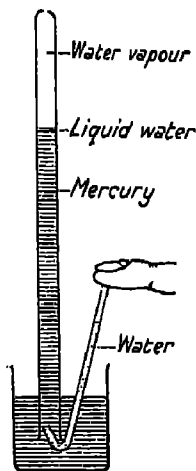
FIG. 261.

Expt. 298. Make a Torricellian barometer as you did in Experiment 39, but this time take care to have a fairly deep trough of mercury. Boil some distilled water until all the dissolved air is driven out. Fill a bent glass tube (Fig. 261) with this water after it has cooled and insert the end of the tube under the barometer tube. Allow a drop of water to rise through the column of mercury to the space above the column, taking care that no air enters. The water evaporates in the space and the level of the mercury falls. Allow a few more drops of water to rise to the top of the column until there is a layer of liquid water on the surface of the mercury. The space is now saturated with water-vapour, and the level of the mercury falls still more.

Now warm the space very gently with a small flame. More water evaporates and the level of the mercury again falls. The vapour pressure increases with a rise of temperature.

These experiments have been performed in a space containing nothing but water-vapour, but the results are also true in air. The air can be saturated with water-vapour, but warm air needs more water to saturate it than cold air. In England the air is seldom saturated whether in summer or winter. The weather is sometimes described as 'muggy' when the air is near its saturation point. It is very unpleasant.

RELATIVE HUMIDITY

We know, in a rough sort of way, what is meant by the dampness of the air, but now that we know that the air can be saturated with water-vapour it is possible to be more accurate in our language. Suppose we find the weight of water in (say) 1,000 c.c. of air at a particular temperature, and then find the weight of water in the same volume of air when it is saturated with water at that temperature. By comparing the two we describe the dampness (humidity) of the air exactly. This is really finding the *relative humidity*, because we are comparing the humidity of the air with its humidity when saturated.

Relative humidity

$$\frac{\text{Wt. of water-vapour present in a given vol.}}{\text{Wt. of water-vapour required to saturate the same vol. at the same temp.}}.$$

It is usually expressed as a percentage by multiplying this ratio by 100. Thus if we say the relative humidity of the air is 50 per cent. we mean that a particular volume of the air contains half as much water-vapour as it would if it were saturated with water-vapour at that temperature.

Another way of expressing the amount of water-vapour in the atmosphere is in terms of the vapour pressure. The ordinary barometer gives the pressure of the atmosphere and this is made up of the pressure due to the air plus the pressure due

to the water-vapour present. These components can be given separately.

For health, the relative humidity should not be too high or too low. In some tropical countries the relative humidity is very high, and Europeans find it very enervating. It is not the heat which causes the discomfort but the high humidity. When there are many people in a badly ventilated room the relative humidity becomes high owing to the water-vapour given out by the lungs. It is this which causes the discomfort, not the decrease of oxygen as is often thought. Occasionally in summer we have a few days during which the relative humidity is very high. Everybody feels lazy and uncomfortable. The perspiration does not evaporate and it runs off the skin as sweat. The inside walls of rooms are covered with condensed water-vapour, and opening all the windows and doors makes no difference because all the air is nearly saturated with water-vapour.

Certain industries require a definite humidity for the work to be done efficiently. The cotton factories of Lancashire are very carefully 'air-conditioned' to keep the air at the right humidity for the process carried on. The climate of Lancashire is supposed to have favoured the beginning of the industry in that county, but with modern air-conditioning it cannot have much importance to-day.

No doubt you have noticed that on some days the water in the pools on the ground after rain evaporates very quickly, but on other days it evaporates slowly. This is noticed both in winter and summer. The rate at which the pools disappear depends on the relative humidity of the atmosphere. When the humidity is high, the pools disappear slowly. Good 'drying days' for the housewife are those on which the relative humidity of the air is low.

There is always some water-vapour even in the air over the driest desert. The sources of all water-vapour in the atmosphere are on the surface of the earth, in oceans, seas, lakes, rivers, the soil, and vegetation, and this is one reason why the lower layers of the atmosphere contain most water-vapour. The higher layers contain very little water-vapour and the highest clouds are mere wisps. An aeroplane flying at about 20,000 ft.

is often visible only by the trail of its exhaust pouring water-vapour into the atmosphere. This immediately condenses, as a very little water is sufficient to saturate the air at that density.

The determination of the relative humidity is important for greenhouses, hospitals, and in meteorology. It can be found directly, but it is a tedious process. A volume of air, say 5 litres, is drawn by means of an aspirator through previously weighed U-tubes containing some substance like calcium chloride or phosphorus pentoxide. The increase in weight gives the weight of water-vapour in the 5 litres of air. The experiment is then repeated, but this time the air is first drawn through water before it is drawn through the U-tubes. This saturates the air with vapour, and the increase in weight of the U-tubes gives the weight of the water-vapour in the 5 litres of air when it is saturated. If the experiment is done while the temperature remains constant, then the first weight divided by the second and then expressed as a percentage, gives the relative humidity of the sample of air. There are, however, simpler but less direct methods as we shall see.

DEW

If a glass of water is taken into a warm room it sometimes happens that moisture forms on the outside of the glass. The moisture is called *dew*. The air in the room is warm and not saturated with water-vapour. In coming into contact with the glass of cold water, the air round the glass is cooled until a temperature is reached at which the air is saturated with the water-vapour present. On further cooling the air can no longer hold all the water as vapour and some is condensed on the sides of the glass. If the air originally is very near the saturation point a small fall in temperature causes the deposition of dew; if, however, the air is a long way from its saturation point it is possible that the fall in temperature caused by the water in the glass will not be sufficient to cause the formation of dew.

The temperature to which air has to be cooled to cause saturation by water-vapour is called the *dew-point*.

Expt. 299. To find the dew-point of the air in a room. To do this Regnault's hygrometer may be employed, but there are many kinds of hygrometers which may be used for this purpose.

At the bottom of two large test-tubes are fitted silver thimbles (Fig. 262). Ether is poured into one test-tube, and the apparatus fitted up as shown in the Figure. When air is drawn into the test-tube and through the ether, the latter evaporates and reduces the tempera-
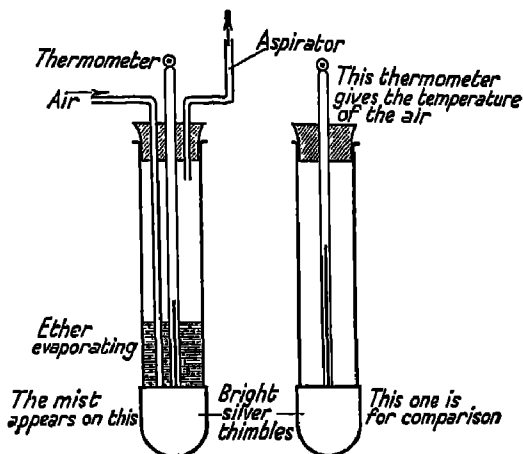
FIG. 262. Regnault's hygrometer.

ture of the silver thimble. In time the temperature falls so low that a mist appears on the thimble. The appearance of the mist is made more evident by comparison with the other thimble. The slightest difference between the two thimbles is sufficient to show the formation of a mist. When the mist appears the thermometer in the ether is read and the aspirator disconnected. As the temperature rises the mist disappears and the temperature is noted when this occurs. The mean of the two temperatures is taken as the dew-point. The readings must be taken with a sheet of glass between the observer and the hygrometer, otherwise the temperature and humidity of his breath will affect the result. The thermometer in the other test-tube gives the temperature of the air.

A simpler method of finding the dew-point, and one which is accurate enough for most purposes, is as follows:

Expt. 300. Put some water into an aluminium can which is highly polished on the outside. Suspend a thermometer in the water from a

stand. Place a sheet of glass between the apparatus and yourself to shield the can from your breath. Add a small piece of ice and stir. Watch the polished surface of the can all the time. When the ice has

melted and reduced the temperature of the water in the can add a little more ice. Continue this process until a film of moisture appears on the polished surface. Take the temperature immediately. Now allow the temperature of the water to rise. Immediately the film of moisture disappears take the temperature. The mean of the two temperatures may be regarded as the dew-point.

THE WET AND DRY BULB HYGROMETER

Expt. 301. Fit up two exactly similar thermometers side by side (Fig. 263). Allow the bulb of one to remain dry in the air of the room. To the bulb of the other fasten a piece of muslin, one end of which trails in a small vessel of water. Note the readings of the two thermometers. The one with the dry bulb gives the higher reading.

The muslin attached to the bulb of one thermometer keeps it wet. The water continually evaporates and thus takes heat from the bulb. This thermometer therefore registers a lower temperature than the other. The rate of evaporation from the wet bulb depends on the amount of water in the air and on the temperature of the air. If the amount of vapour in the air is small, the rate of evaporation is great and there is a large difference between the temperatures registered by the two thermometers. By noting this difference we have an indication of the amount of water-vapour in the air. Such an hygrometer may be used for finding the relative humidity if tables are previously made showing the relation between the relative humidity, the temperature of the air, and the difference between the readings of the two thermometers, e.g. the relative humidity is 68 per cent. when the temperature of the air is 15° C. and the wet bulb thermometer reads 12° C.

FIG. 263. Wet and dry bulb hygrometer.

THE FORMATION OF DEW ON THE GROUND

We often find dew on the ground in the early morning. The cold earth cools the air near it below the dew-point and dew is deposited. During the night-time when there is a clear sky and the air is still, the earth radiates heat. It is after such nights that a heavy deposit of dew is found.

All the water we see on the grass in the morning is not dew. Much of the water is given out by the grass. Usually when a plant transpires water-vapour is given off. If, however, the relative humidity of the air is high, the rate of transpiration is low, but sometimes, especially in the case of grass and similar plants, the water leaves the plant as a liquid. This is not dew.

If the dew-point is below 0° C. and the air is cooled to the dew-point, then *hoar frost* is formed instead of dew, solid crystals being obtained instead of drops of water.

On a cloudy night dew or hoar frost is seldom formed. It is on cold, dry, clear nights we expect frost. The atmosphere acts as a blanket to the earth, helping it to keep its heat. The presence of water-vapour and cloud increases the efficiency of the atmosphere as a blanket. The radiation from the earth is either absorbed by the clouds or reflected to the earth and most of the heat is kept in the lower layers. On dry, clear, cloudless nights there is less to absorb the earth's radiation and much of it escapes to the upper layers.

FOG AND MIST

There is really no distinction between a mist and a fog. A dense mist is a fog. Fogs are formed in various ways. They are often formed in valleys where the air is still and the earth cold. Comparatively warm and moist air coming in contact with this ground would cause a mist or fog. In some districts a fall in the wind very often causes a fog. On the other hand, wind can cause a fog by bringing warm damp air into contact with cold land.

The cold earth cools the air to the dew-point and if particles of dust, such as the products of the combustion of fuel, are present in the air, the water-vapour condenses into tiny droplets

each having one of these particles, known as hygroscopic particles, for a nucleus. The droplets are so tiny that they remain suspended in the air. The heavy fogs of our large cities are formed in the same way as other fogs, but once formed they prevent the escape of the solid particles in the air and the water collects round these particles also. The smoke of cities does not cause the fogs but it makes them much more persistent and unpleasant. Any person walking through a city fog will be left in no doubt that much solid matter is contained in the fog. The face and clothing are much soiled.

Sea fogs are formed by winds blowing warm, moist air over the surface of the cold sea.

CLOUDS, RAIN, AND SNOW

The air is usually rising or falling, sometimes slowly, sometimes in gusts. When air rises it cools, because it expands owing to the decreased pressure and also because it receives less of the long-wave radiation from the earth. This cooling continues until at last the air has been cooled to the dew-point and is now saturated with water-vapour. Some vapour now condenses round hygroscopic particles to minute drops and these continue to ascend, carried up by the rising air. As they rise the drops increase in size and there is a greater tendency for them to fall, but the air is still rising and prevents them from falling. It is these innumerable drops of water which constitute a *cloud*.

As the air column continues to rise, the drops increase in size until they are too big to be supported by the up-rising air. They then fall as *rain*.

If the condensation takes place when the temperature is below o° C. then solid particles are formed and these fall as *snow*. Snow-flakes are bundles of interlocked crystals. As they pass through warmer layers they may be melted and fall as rain.

HAIL

The formation of hail depends on the fact that it is possible to *super-cool* water. Pure water may be cooled under certain conditions to a temperature as low as $-20°$ C. before it is frozen. Consider moist air rising. When it gets to the cloud level, condensation takes place, but the minute drops of water so formed continue to ascend with the rising column of air. At a much greater height the temperature falls below $0°$ C., but the drops do not freeze but are super-cooled. Continuing to ascend, they reach a height where snow crystals are formed. These rapidly grow and begin to fall. In their descent they pass through the ascending column of super-cooled water which freezes round the snow-flakes imprisoning some air. Thus the hail-stone at this stage has a coating of soft white ice. The stones continue their descent and meet water at $0°$ C. which wets the surface of the soft white stones and then freezes on them as clear ice. In this condition it often falls to the ground as *hail*.

Examine the next hail-stones you find. Cut some in two and try to find the concentric layers of clear ice and soft white ice. Large stones may show several such concentric layers. This is because they have been carried up and down several times in their journey to the earth.

QUESTIONS

1. Describe a simple experiment you could perform to show that there is water in the atmosphere. What everyday observations lead us to suppose that there must be water in the atmosphere? How could you obtain a specimen of dry air?

2. Explain carefully what is meant by the terms: (a) Relative humidity, (b) Dew-point. Under what conditions do clothes, hung in the air, dry quickly? What influence has the humidity of the air on the comfort and health of man?

3. Describe, with the aid of a sketch, the Wet and Dry bulb hygrometer. On what principle does it work, and what is its use?

4. How, and under what conditions, is dew formed on the ground? What is hoar frost? Why is hoar frost seldom formed on cloudy nights?

5. In what state is the water in a cloud? Describe one way in which clouds may be formed. What keeps them floating in the air?

6. Describe the formation of (*a*) hail, and (*b*) snow.

XXX

MOVEMENTS OF THE ATMOSPHERE: WEATHER

EVERYONE knows that the atmosphere is almost constantly in motion. The horizontal movements of the atmosphere, or *winds*, are obvious, but the *vertical movements* are not so obvious. They are, however, as important as the winds in their influence on the *weather*.

The study of the weather has always been important to the ordinary man but especially so to the sailor and the farmer, whose livings depend directly on it. To the airman it is of vital importance because he cannot make a journey from one place to another unless he has some idea of the strength and direction of the wind. If he is likely to meet bad weather on his journey he wants to know as much about it as possible so that he can be prepared for it, or, by taking a different route, avoid it.

We now know much more about the science of the weather (*meteorology*) than we did a few years ago. As knowledge increases, forecasting becomes more accurate, but even to-day there is no certainty about a weather forecast, because the factors which influence the weather are very complex, are not fully understood, and a great amount of information about conditions over a very wide area is necessary before any forecast can be attempted. Even with the widespread organization which exists for collecting this information it is not possible to obtain all the necessary facts.

Weather science is a fascinating study which many pursue for its own sake. In this book we have already dealt with many of the basic principles on which the science rests and in our study of air pressure, heat, and water in the atmosphere we have often dealt with weather topics. Before leaving the subject, for in a book of this character we cannot pretend to do more than introduce it, we shall deal in an elementary way with the movements of the atmosphere which have such a great

influence on the weather. It may even be said that they con-
stitute the weather.
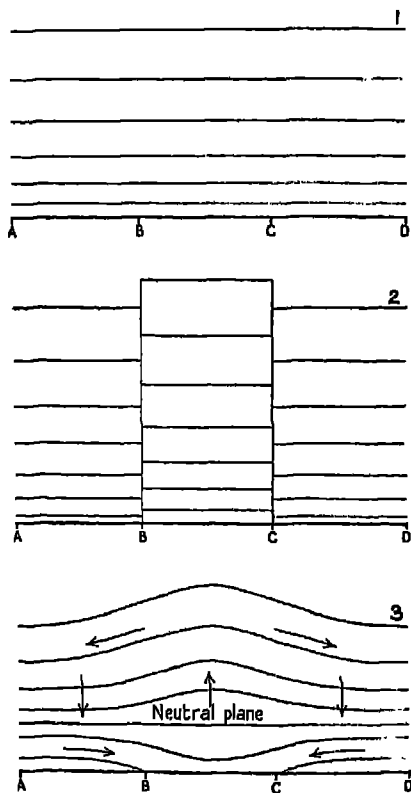
FIG. 264. Effect of heating on pressure in the
atmosphere.

DIFFERENCE OF PRESSURE AND UNEQUAL HEATING OF
THE ATMOSPHERE

If the pressure of the atmosphere is everywhere the same the
air is at rest. The air is set in motion by the differences of

pressures in the atmosphere. The air travels from a position of high pressure to one of lower pressure tending to make the pressures equal. We saw this was so when we considered Land and Sea Breezes. Let us consider the question in more detail.[1]

Fig. 264(1) represents a section of the atmosphere *ABCD* representing the earth. There is no movement in the air and the pressure at *A, B, C, D,* is the same. Now suppose the air just above *BC* is heated (Fig. 264(2)). It expands and rises. Imagine that the walls of the column of air are air-tight. The pressure on *BC* is not changed because there is the same weight of air above it as before. Now suppose the air-tight walls to be removed. Air flows outwards from the top of the column *BC*, some piling on column *AB* and some on column *CD* (Fig. 264(3)). The pressure on *AB* (and *CD*) is now increased because air has been added to it. The pressure over *BC* has been reduced because it has lost air. Air now flows from the lower layers of *AB* and *CD* inwards to *BC*, thus tending to restore the pressure. If the heating of *BC* is maintained the circulation continues, air rising above *BC*, flowing outwards, descending over *AB* and *CD* and flowing inwards to *BC*.

THE UNEQUAL HEATING OF THE ATMOSPHERE

Fig. 265 is a map of the world on which is shown the isotherms for the month of January. An *isotherm* is a line drawn through all places on the earth which have the same temperature at the same time. Observations and records are kept at many stations all over the world. The temperature of the air is taken and reduced to what it would be at sea-level, allowing about 1° F. for every 300 ft. of altitude. A line joining all points of equal temperature at the time of observation is then drawn. Now there are many factors controlling the temperature of the air, but there are three very important ones: (1) Latitude, (2) Altitude, and (3) Distance from the sea. If latitude were the only factor, the isotherms produced would be parallel to lines of latitude. If altitude were the only factor the isotherms would

[1] We follow here a Figure and explanation given by W. G. Kendrew in *Weather, an Introductory Meteorology.* O.U.P.
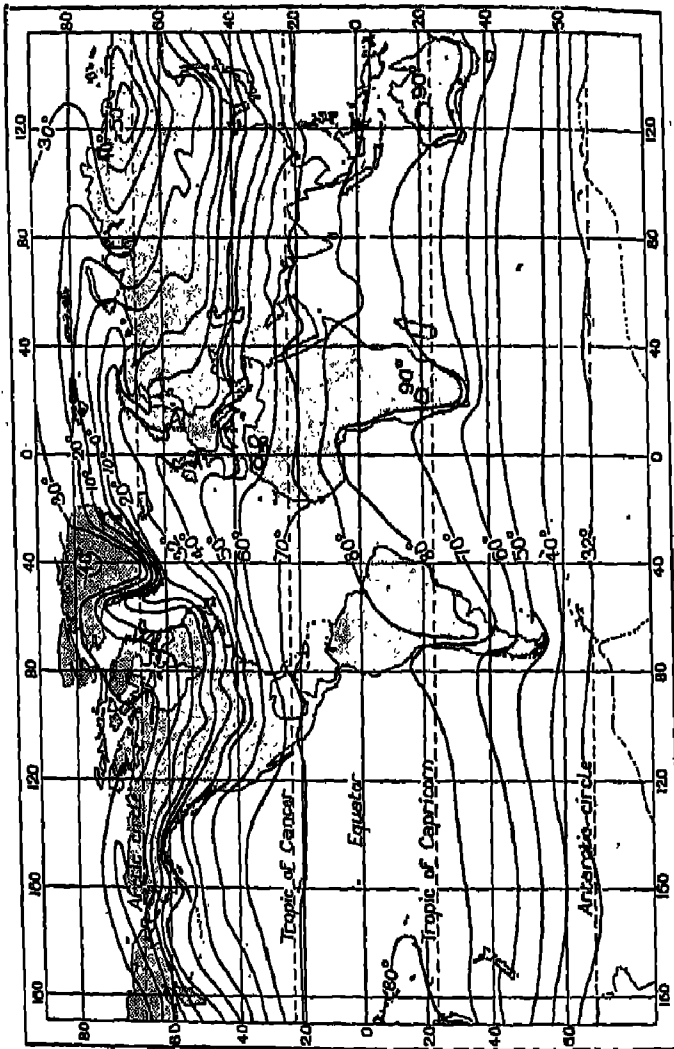
FIG. 265. January isotherms.

follow the contours of the land. 'Reduced' isotherms eliminate the effects of altitude and show only the effects of latitude and distance from the sea. The effects of latitude are shown by the east to west direction of the contours, hence any departure from the east to west direction is due largely to distance from the sea.

It will be seen from the map that the isotherms run very roughly parallel to the equator, the temperature decreasing from the region of the equator towards the poles. The iso-
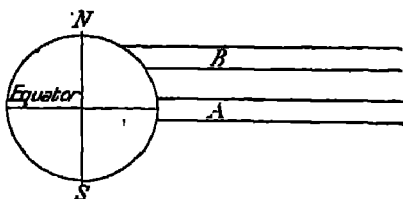
FIG. 266. The sun's rays heating the earth.

therms are much more irregular over the land masses than over the sea. It will also be noted that the southern hemisphere is warmer than the northern in this month. The July isotherms (not shown on this map) show the same general tendency to be parallel to the equator, but the southern hemisphere is now colder than the northern.

The sun is so far from the earth that all the rays reaching the earth from the sun may be considered to be parallel. Fig. 266 shows how this fact brings about unequal heating on the earth. The sun is directly overhead at the equator (when the earth is in this position) and beam A which strikes the earth at the equator covers a smaller area of the surface of the earth than does an exactly similar beam B which strikes the earth in the latitude of London. This means that the same amount of heat is spread over a larger area near London. Thus the earth in this region receives less heat per square mile than does the earth at the equator.

CII.

FIG. 267. January isobars.

THE DISTRIBUTION OF PRESSURE IN THE ATMOSPHERE
NEAR THE SURFACE OF THE EARTH

This unequal heating of the earth's surface resulting in the unequal heating of the atmosphere brings about unequal pressures, as we have seen. Fig. 267 is a map of the world showing the distribution of pressure over the earth in January. An *isobar* is a line drawn through all points on the earth's surface which have the same atmospheric pressure at the same time. The pressures, like the temperatures for isotherms, are first reduced to what they would be at sea-level, otherwise the isobars would follow the contour lines on the map, as pressure decreases with height. The isobars on our map show the mean pressure in inches for January. Note that pressure is highest in the sub-tropical zones of both hemispheres. These regions of high pressure are known as the sub-tropical anti-cyclones. The pressure is less in the equatorial region. The pressure also falls north of the sub-tropical anti-cyclones in the northern hemisphere and south of them in the southern hemisphere. The pressure rises a little as one approaches the poles in both hemispheres.

THE DIRECTION OF THE WIND

If we consider only the difference in pressure between two points on the earth's surface we should expect the wind to blow from the region of high pressure to that of the low. But there are other factors to take into consideration. The rotation of the earth deflects the wind towards the right of the direction in which it is moving in the northern hemisphere and towards the left in the southern hemisphere. The deflection is greatest at the poles and decreases towards the equator where there is none. The result is that the wind is deflected until it flows along the isobars and not at right-angles to them. The friction between the moving air and the surface of the earth, however, slows the wind down and causes it to travel at an angle to the isobars. At a height of about 1,500 ft. the friction is not felt, so the wind at that height flows along the isobars. These, however, are not the same as those at the surface. No

doubt you have often noticed clouds being blown in a different direction from the wind direction near the ground. Friction

FIG. 268.

is least over the sea and greatest over uneven land such as hilly country, towns, and wooded country.

BUYS BALLOT, a Dutch meteorologist, formulated a law showing the relation between pressure and wind direction. It is known as *Buys Ballot's law*: 'Stand with your back to the wind, then in the northern hemisphere the pressure on your left is lower than that on your right. In the southern hemisphere it is higher on your left than on your right.' This means that in the northern hemisphere the wind travels clockwise around a high-pressure system and anti-clockwise around a low-pressure system. The reverse is true in the southern hemisphere. The wind, however, is never perfectly steady. It is always *veering* (altering its direction clockwise) or *backing* (altering its direction anti-clockwise) a little (Fig. 268).

THE VELOCITY OF THE WIND

The velocity of the wind is measured by an instrument called an anemometer of which there are various types. The strength of the wind varies from *calm* (less than 1 m.p.h.) to that of a *hurricane* (above 75 m.p.h.). Wind above 60 m.p.h. is very rarely experienced in Great Britain and when it occurs widespread damage is caused. A wind between 25 and 31 m.p.h. is called a *strong breeze* and it is difficult to use an umbrella in it. The winds have been classified according to their velocities (strengths) and, in the Beaufort Scale of Wind Force, each is given a number and description so that anyone reporting a wind need give only the number, and this fully describes it. For example, if a wind is given as Beaufort No. 4 its velocity is between 13 and 18 m.p.h. It is called a *moderate breeze* and is recognized by its raising dust and loose paper and moving

small branches of trees. The Beaufort Scale is given in all books on Weather.

The wind is never perfectly steady. It blows in *gusts*. These are caused by obstructions on the ground; hence gustiness is greater over the land than over the sea. For the same reason gustiness is less at a height where the effect of the obstructions is not felt. It is greater with a strong wind than light. These gusts may be dangerous to a pilot when landing or taking off, but the gusts caused by large permanent features such as hills, trees, and buildings near an aerodrome are more or less regular and the pilot can learn them and allow for them.

VERTICAL MOVEMENT OF THE AIR

We have already seen that convection is one very important cause of vertical movements in the atmosphere. Another is the relief of the land. When a wind comes in contact with mountains the air is forced upwards. As a result it expands and therefore cools and often becomes saturated with water-vapour. Clouds are formed and rain may fall. Rain formed in this way is known as *orographic rain* or *relief rain*. The climate of some districts is determined by orographic rain caused by a prevailing warm moisture-laden wind striking the face of high mountains.

Air is also forced to rise when cold and warm currents of air meet. The warm air is forced upwards, the cold air flowing underneath it. We see a similar thing happen when oil and water meet. The lighter oil is forced on top of the denser water. Warm air thus forced up becomes cooler, clouds are formed and often rain falls. This is called *frontal rain*. When very cold air meets warm air the vertical currents formed are sometimes very violent and are accompanied by thunder and lightning. Some of the upward currents of a thunderstorm thus caused attain a velocity of 20 m.p.h.

Bumps, well known to the airman, are vertical movements of the air caused by unequal radiation of heat from the surface of the earth and in other ways. They are commonly met when an aeroplane reaches the air over the land after coming from over the sea. The radiation of heat from the land is greater than

that over the sea and this causes a local rising current where the land begins. A wind coming from the sea is sometimes forced up on meeting the land and this also causes bumps. Bumps are not common over the sea because its surface is uniform as regards the radiation of heat, but over land they are common and are sometimes experienced by an aeroplane over the borders of grassland and bare soil. The latter is usually warmer than the grassland and radiates more heat to the air immediately above it.

Downward movements of the air are not so noticeable as upward movements because they are usually spread over a greater area and are, therefore, less violent. As the air descends its volume decreases (owing to the increased pressure) and, therefore, it becomes warmer. This is not often noticed as some of the heat is again lost as latent heat in vaporizing cloud, and there is little final change in temperature, but in some parts of the world, notably in the Alps and the Rocky Mountains, there are well-known warm winds coming down the sides of the mountains to the valleys.

Usually, rising currents are cloudy. We have seen why. The type of cloud produced by rising currents is the 'cauliflower' or 'heap' type, *cumulus cloud*. They are often very deep and sometimes reach a height of 20,000 ft. The appearance of cumulus cloud is taken as a sign of ascending currents and glider pilots look for them so that they can take advantage of ascending currents and so keep in the air as long as possible.

Layer or *stratus* clouds have no pronounced vertical structure. They are common at sunset in fine weather. They usually indicate that there is little vertical movement of the air, and these stable conditions usually accompany fine weather.

Descending currents are usually free from cloud because the air is contracting and so becoming warmer. When a cloud begins to descend it often disappears because the air can contain more moisture without becoming saturated.

THE TROPOSPHERE AND STRATOSPHERE

We saw (p. 442) that the temperature of the air decreases with height, the dry adiabatic lapse rate being 1° C. for every 100 metres rise; but the water-vapour in the atmosphere has an influence on the rate of fall of temperature with increased height. Rising air is soon cooled to the dew-point and water-vapour condenses to water. The latent heat of vaporization given out during the process warms the mass of air. Thus when moist air rises we have two processes going on at the same time, one causing loss of heat due to the air expanding and the other bringing about gain of heat from the latent heat of vaporization. The heat gained in the second process accounts for a rise in temperature of about 0·5° C. for every 100 metres increase in altitude. The difference between this and the dry adiabatic lapse rate (1° C. per 100 metres) is called the *saturated adiabatic lapse rate* and its value in the temperate regions is 0·5° C. for every 100 metres rise. The actual lapse rate may vary considerably from this on any particular day.

The temperature of the air in the upper regions is found by sending up small hydrogen balloons, each carrying a thermometer. In this way the mean temperature of the air at various levels to a very great height is known. We should expect the temperature to continue to fall with increase in altitude, but it does not. At a height of about 6 miles in Great Britain the temperature stops falling and remains steady at about —51° C. no matter how high the balloon ascends. Sometimes there is even a slight rise. The layer of the atmosphere where the temperature ceases to fall is called the *tropopause*. All the atmosphere below the tropopause is call the *troposphere* and all above it the *stratosphere*. At the equator the tropopause is about 11 miles high; at the poles it is only 4 miles high.

WIND ESTIMATION BY AIRMEN

The pilot of an aircraft must know the direction and speed of the wind in which he is flying if he is to set a proper course (p. 521). Before he leaves his station he obtains all the

information he can about the speed and direction of the wind he is likely to meet on his journey. He can also obtain information during his journey by radio communication with his own and other stations. He cannot assume that the strength and direction of the wind will remain the same. Indeed, if the journey is long he may expect all kinds of changes in the weather. Sometimes he must rely on his own observations. By practice he can make a fair estimation of the speed of the wind he is travelling in by noticing the speed of the shadows of clouds on the ground. These clouds should be at the same height as the aeroplane, for we have seen that the speed and direction of the wind sometimes varies with the height. He can also find the direction of the wind by comparing the direction of the moving shadows with the compass in his aircraft. When travelling over the sea he notes the appearance of the surface and this sometimes helps him to find the strength and direction of the surface wind. When landing it is important for a pilot to know about the wind near the surface at his landing-ground. He notes smoke from chimneys, flying flags, the behaviour of trees, and even the swaying of grass. From these observations an experienced pilot can estimate the velocity of the wind within a few miles per hour.

Wind Velocity and the Navigation of an Aeroplane

When we speak of the velocity of an aeroplane (or anything else) we include both its speed and direction, e.g. we can say that the velocity of an aeroplane is 300 m.p.h. due north or perhaps 250 knots (nautical miles per hour) at a bearing of 095°. (Degrees are reckoned clockwise from true north, there being 360° in one complete revolution.)

Velocity is defined as the rate of change of position in a given direction.

When an aeroplane is in the air its velocity with respect to the ground depends on (1) its velocity through the air, and (2) the wind velocity. The aeroplane is carried with the air at the velocity of the air and at the same time it is thrust through the air by its airscrew. Imagine a ship travelling on the sea at a speed

of 20 knots due north. Suppose the tide is flowing at a speed of 5 knots due south. Then the speed of the ship due north with respect to the earth is 15 knots. The ship goes bodily with the tide while at the same time it goes through the water pushed by its propeller. If the ship goes through the water at only 5 knots due north and the tide is running at 5 knots due south, then the ship is making no progress towards the north. With respect to the earth it is not moving.

It is the same with an aeroplane in the air. It may be travelling through the air at 400 m.p.h., but if it is meeting a head-on wind of 50 m.p.h. then its speed over the ground is only 350 m.p.h. We must, therefore, distinguish between *air speed* and *ground speed*.

The *wind velocity* is the speed and direction of the wind over the earth. Suppose an aeroplane is heading due east, that is, its nose is pointing east and its longitudinal axis is on a line running due east. We say that its course is due east (or, 090°). The *course* is the angle between a true north and south line (meridian) and the longitudinal axis of the aeroplane.

An aeroplane may keep such a course and yet not travel due east because the wind may be blowing from the north and the aeroplane would be carried towards the south. The direction it would travel would be one somewhere between east and south depending on the relative speeds of the aeroplane and the wind. This direction is called the *track*. It is the angle between a true north and south line (meridian) and the actual path of the aeroplane relative to the ground.

THE PARALLELOGRAM OF VELOCITIES

Just as on p. 374 we represented forces (in magnitude and direction) by straight lines, so can we represent velocities, and the same argument by which we established the parallelogram of forces can be used to establish the truth of the parallelogram of velocities. Velocity involves both speed and direction. A line representing a velocity represents the speed by its length, and the direction by reference to a meridian.

Let us consider an example:

In Fig. 269 *NOS* is a meridian. *OC* represents by its length the air speed of an aeroplane (in this case 200 m.p.h.) and by its direction the course (060°). In the same way *O W* represents the wind velocity, the speed (50 m.p.h.) is represented by its length drawn to the same scale as *OC*, and its direction by the angle it makes with the meridian (160°). Now if we
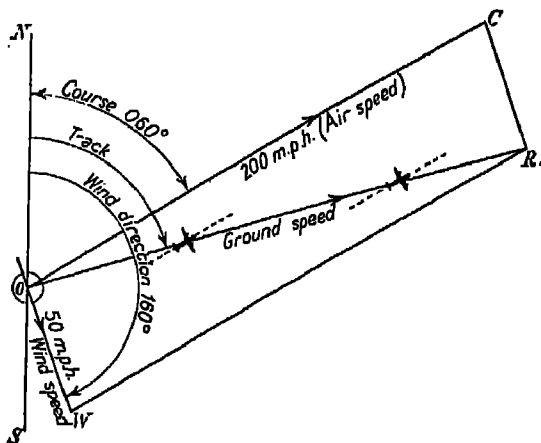
Fio. 269. Parallelogram of velocities.

draw *WR* parallel to *OC* and *CR* parallel to *O W* we have a parallelogram *OCRW* the diagonal of which, *OR*, represents the resultant velocity with respect to the ground. The length of *OR* represents the speed of the aeroplane over the ground (*ground speed*) and the angle which *OR* makes with the meridian *NO* is the track over the ground. Both the ground speed and the track may be obtained by measurement or calculation. It will be noted that although the aeroplane travels in the direction *OR* over the ground it is always pointing in the direction of its course. No doubt you have sometimes seen an aeroplane in a strong wind not travelling exactly in the direction it is pointing but drifting a little to one side. The angle *COR* is called the *drift*.

THE TRIANGLE OF VELOCITIES

It will be seen from the figure that we do not need all the parallelogram to find the track and the ground speed. The triangle *O WR* is sufficient because *WR* is equal and parallel to
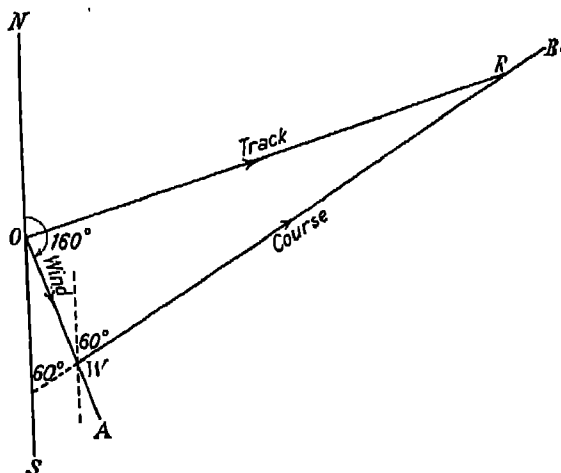
FIG. 270. Triangle of velocities.

OC. In aircraft navigation the parallelogram is never used as it is quicker to use the triangle of velocities.

To find graphically the ground speed and the track when given the course, air speed, and wind velocity using the triangle of velocities. We take the same values as before (Fig. 270).

Draw the meridian *NOS*.

With a protractor construct the angle *NOA*, 160°.

Choose a unit and make *OW* represent 50 m.p.h.

At *W* (not at *O*) draw *WB* making an angle of 60° with the meridian.

With the same unit used in *OW* cut off *WR* to represent 200 m.p.h. Join *OR*.

Measure (or calculate) (1) *OR* for the ground speed, (2) the angle *NOR* for the track.

With a little thought you should now be able to solve graphically (or by calculation if you can solve a triangle) the problems given on p. 524. They are the kind of problems the navigator of an aeroplane has to solve. Remember to put an arrow on each line to indicate the direction. You will find that all the arrows do not go in the same direction round the triangle. The arrow on the resultant velocity is always in the opposite direction to that of the two component velocities. Use common sense as a check on your work and, if you get puzzled, try to solve the problem again with the parallelogram of velocities, and then go back to the triangle to see where you were wrong.

QUESTIONS

1. Describe how the atmosphere is heated and state the causes of the unequal heating. How does this unequal heating lead to differences of pressure in the atmosphere?

2. What is an isobar? Describe, with the aid of a sketch-map, the distribution of pressure in the atmosphere near the surface of the earth.

3. State Buys Ballot's Law. What causes influence the direction of the wind? Explain the meaning of (a) veering, (b) backing.

4. Give three causes of vertical movements of the atmosphere. In what circumstances may an airman experience 'bumps'?

5. Compare and contrast (a) rising currents, and (b) downward currents of the atmosphere with respect to (i) temperature change, (ii) cloud formation.

6. Distinguish between orographic rain and frontal rain.

7. Explain the terms: Dry adiabatic lapse rate, Saturated adiabatic lapse rate, Tropopause.

8. In what ways does an airman, out of communication with the ground, estimate the velocity of the wind?

9. An aeroplane with an air speed of 300 knots is flying due east. A wind of 20 knots is blowing from the north-east. By means of a scale drawing find the track and ground speed.

10. An aeroplane requires to make a track due west and its air speed is 250 knots. There is a 25-knot wind blowing from the north-west. Find, by means of a geometrical construction, what course it must steer and what will be the ground speed.

1. A cyclist weighing 150 lb. starts from rest and coasts down a hill half a mile long with a gradient of 1 in 10 (i.e. the road falls 1 ft. for every 10 ft. measured down the hill). (i) Neglecting friction and wind resistance, calculate his kinetic energy when he reaches the bottom of the hill. (ii) What is the name given to the type of energy from which this kinetic energy is derived? (iii) Give an example of the reverse transformation of energy, viz. the conversion of kinetic energy to the type of energy referred to in (ii). [N.J.B.]

2. Why is a lever classed as a machine? Give an account of the various orders of lever in common use.

A person using a pair of cutting pliers exerts a force of 8 lb. weight in order to cut a wire. If the wire is $\frac{1}{2}$ in. away from the pivot and the force is applied 6 in. from the pivot, calculate the resistance offered by the wire. [C.W.B.]

3. What is formed when carbon dioxide and water are shaken together in a large bottle? Compare this with what happens to the carbon dioxide and water taken in by a green plant in sunlight. Explain the use of solid carbon dioxide in ice-cream carts. [Lond.]

4. How is work measured in physics? What is the horse-power of an engine capable of pulling a train on the level at a mile a minute against a frictional resistance of 1 ton? (1 horse-power = 550 ft.lb. per sec.) [O. & C.]

5. Give an account of an experiment to demonstrate convection currents in air. Describe the heating and ventilating system, in any building you know, by means of a clear labelled diagram together with a short verbal explanation. [Lond.]

6. What do you understand by a fuel? Name one solid, one liquid, and one gaseous fuel. Draw a labelled diagram of the bunsen burner. State the two distinct kinds of flame it can produce and explain why one of these is unsuitable for general laboratory work. Explain, briefly, how you would show that there is a region of unburnt gas in the bunsen flame. [C.W.B.]

7. A car weighs 3,000 lb. and its weight is distributed evenly between the four wheels. Calculate the force that must be exerted and the work done in lifting one wheel 1 in. off the ground, using a jack with a mechanical advantage of 150 and an efficiency of 40 per cent. [O. & C.]

8. (i) What conditions must be satisfied if three coplanar forces acting at a point are to be in equilibrium?

A buoy, volume 100 cu. ft. and weight 300 lb., is attached by a

rope to the sea bottom. A strongly running tide exerts on the buoy a horizontal force of 200 lb. weight and one-tenth of the volume of the buoy is submerged. Determine by means of a careful drawing the tension in the mooring rope under these conditions. (1 cu. ft. of sea-water weighs 64 lb.)

(ii) Name and sketch, but do not describe, one mechanical device in which a weight is supported by the application of two other forces.

[N.J.B.]

9. Choose four examples of levers you use in every-day life and describe, with the help of diagrams, how they confer mechanical advantage on the user.

An object weighing $3\frac{1}{2}$ cwt. has to be raised from the ground by a man who can exert a force of 1 cwt. Describe an arrangement of pulleys, hanging from a beam, which makes this possible. What is the velocity ratio of the apparatus you describe? [Lond.]

10. Explain what is meant by uniform acceleration. Calculate the time that would be taken for a bomb to fall 10,000 ft. if air resistance could be neglected. ($g = 32$ ft. per sec. per sec.)

[O. & C.]

11. How would you distinguish between: (a) two yellow 'metals', one gold and the other brass, (b) two white powders, chalk and starch, (c) cupric oxide and manganese dioxide, (e) dilute sulphuric acid and dilute hydrochloric acid? [Lond.]

12. A fixed and movable pulley block are employed in such a manner that the velocity ratio is 4 when a vertical downward effort is used to move a load. Draw a sketch of the arrangement. The mechanical advantage is found to be 3; what is the efficiency of the machine? If the load is 96 lb., what effort is needed? What is the principal factor reducing the efficiency of the machine? [Lond.]

13. Why does a ship float even though it is largely made of a heavy substance like steel?

If a ship of 10,000 tons had a uniform cross-section of 25,000 sq. ft. what depth of the vessel would be submerged in sea-water whose density is 64 lb. per cu. ft.?

14. Describe in detail an experiment which you would perform to show that the specific heat of water differs from that of alcohol.

[N.J.B.]

15. What is the difference between heat and temperature? Describe any one experiment to illustrate the distinction. If 1,000 British Thermal Units were given to 100 lb. of water at freezing-point, what would be the final temperature of the water (a) in degrees Centigrade, (b) in degrees Fahrenheit? [Lond.]

16. Explain carefully what happens when a pond of water originally at 12° C. is cooled down until a thick layer of ice is formed. Illustrate your answer with diagrams showing the temperature of different parts of the pond as the cooling proceeds.   [O. & C.]

17. Describe, giving one example in each case, the methods by which heat is transferred. Show that in a thermos flask transference of heat by all methods is reduced to a minimum.   [C.W.B.]

18. What do you understand by the terms (a) kinetic energy, (b) potential energy, (c) the calorific value as a fuel?

Give a clear labelled diagram of the essential parts of either (i) a motor-car engine, or (ii) a steam locomotive engine and describe how the energy liberated by the combustion of the fuel is transmitted to the driving wheels. Is the internal- or the external-combustion engine more efficient? How is energy wasted in the engine you have described?   [Lond.]

19. Describe, with the aid of a diagram, a clinical thermometer. What peculiarities are there in its construction?

Explain how two thermometers can be used to estimate the relative humidity of the atmosphere. How does the amount of water-vapour in the atmosphere effect physiological processes in plants and animals?   [Lond.]

20. Explain shortly: (a) Why the spout of a teapot reaches up at least as high as the lid. (b) Why there is a risk of cracking a thick glass tumbler if very hot water is poured quickly into it. (c) Why water-pipes sometimes burst in winter. (d) How a fire in an open grate helps to ventilate a room.   [C.W.B.]

21. What is meant by the expressions 'transformation of energy' and 'conservation of energy'? Briefly explain how the following act as transformers of energy: (a) an electric lamp, (b) a dynamo, (c) chlorophyll. How does the heat produced by an electric current, passing through a fixed resistance, depend upon the strength of the current?   [N.J.B.]

22. Explain the devices you would make use of (a) to keep soup hot for as long as possible, and (b) to prevent oneself from becoming over-heated in the tropics.   [O. & C.]

23. Write the equation for any chemical reaction which is (i) a reduction, (ii) the action of an acid on an insoluble base, (iii) the process resulting in the liberation of hydrogen.   [N.J.B.]

24. Name the three processes you would apply in succession to obtain a pure crystalline specimen of sodium carbonate from a dirty specimen of washing soda. The crystals of sodium carbonate will be efflorescent, i.e. they will in time become covered with a white powder

and thus cease to look clear. What is the cause of this behaviour?

[N.J.B.]

25. How would you prepare (*a*) a crystallized specimen of zinc sulphate from granulated zinc, (*b*) soap from mutton fat?   [C.W.B.]

26. Describe the preparation of a solution of hydrochloric acid from common salt. How could you obtain from such a solution (*a*) chlorine, (*b*) hydrogen?   [O. & C.]

27. Explain *three* of the following: (*a*) lime-water will soften temporarily hard water; (*b*) moist chlorine is a bleaching agent; (*c*) lead oxide is converted into lead when heated on a piece of charcoal; (*d*) quicklime can be used to dry gases.   [O. & C.]

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